Properties of Acids & Bases

Your notes

Properties of Acids

- Acids have pH values of below 7, have a sour taste (when edible) and are corrosive
- Acids are substances that can neutralise a base, forming a salt and water
- When **acids** are added to water, they form positively charged **hydrogen ions** (H^+)
- The presence of **H^+** ions is what makes a solution acidic

Example: Hydrochloric Acid

$$HCl\,(aq) \rightarrow H^+\,(aq) + Cl^-\,(aq)$$

Typical reactions of acids

Acids and metals

- Only metals **above hydrogen** in the reactivity series will react with dilute acids.
- When acids react with metals they form a salt and hydrogen gas:

Acid + Metal → Salt + Hydrogen

- The name of the salt is related to the name of the acid used, as it depends on the **anion** within the acid.
- Examples of the names of salts from specific acids and metals are:

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride and hydrogen	$Mg + 2HCl \longrightarrow MgCl_2 + H_2$
Sulfuric acid	Magnesium sulfate and hydrogen	$Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$
Nitric acid	Magnesium nitrate and hydrogen	$Mg + 2HNO_3 \longrightarrow Mg(NO_3)_2 + H_2$

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Acids with bases

- Metal **oxides** and metal **hydroxides** (**alkalis**) can act as **bases**
- When they react with acid, a **neutralisation** reaction occurs
- In all acid-base neutralisation reactions, **salt** and **water** are produced

Acid + Base → Salt + Water

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Your notes

- Examples of reactions between acids and bases:

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride and water	$Mg(OH)_2 + 2HCl \longrightarrow MgCl_2 + 2H_2O$
Sulfuric acid	Magnesium sulfate and water	$MgO + H_2SO_4 \longrightarrow MgSO_4 + H_2O$
Nitric acid	Magnesium nitrate and water	$Mg(OH)_2 + 2HNO_3 \longrightarrow Mg(NO_3)_2 + H_2O$

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Acids with metal carbonates

- Acids will react with metal carbonates to form the corresponding metal salt, carbon dioxide and water:

Acid + Metal Carbonate → Salt + Carbon Dioxide + Water

- Examples of reactions between acids and carbonates:

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride, carbon dioxide and water	$MgCO_3 + 2HCl \longrightarrow MgCl_2 + CO_2 + H_2O$
Sulfuric acid	Magnesium sulfate, carbon dioxide and water	$MgCO_3 + H_2SO_4 \longrightarrow MgSO_4 + CO_2 + H_2O$
Nitric acid	Magnesium nitrate, carbon dioxide and water	$MgCO_3 + 2HNO_3 \longrightarrow Mg(NO_3)_2 + CO_2 + H_2O$

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Indicators

- Two colour indicators are used to distinguish between acids and alkalis
- Many plants contain substances that can act as indicators and the most common one is **litmus** which is extracted from lichens
- Synthetic indicators are organic compounds that are sensitive to changes in acidity and appear different colours in acids and alkalis
- **Thymolphthalein** and **methyl orange** are synthetic indicators frequently used in acid-alkali titrations

Two Colour Indicators Table

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Indicator	Colour in acid	Colour in alkali
Litmus	red	blue
Thymolphthalein	colourless	blue
Methyl orange	red	yellow

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- Synthetic indicators are used to show the **endpoint** in titrations as they have a very sharp change of colour when an acid has been neutralised by alkali and vice-versa
- Litmus is not suitable for titrations as the colour change is not sharp and it goes through a purple transition colour in neutral solutions making it difficult to determine an endpoint
- Litmus is very useful as an indicator paper and comes in red and blue versions, for dipping into solutions or testing gases

Your notes

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Properties of Bases & Alkalis

Your notes

- Bases have pH values of above 7
- A base which is water-soluble is referred to as an **alkali**
- In basic (alkaline) conditions red litmus paper turns **blue**, methyl orange indicator turns **yellow** and thymolphthalein indicator turns **blue**
- Bases are substances which can neutralise an acid, forming a salt and water
- Bases are usually **oxides** or **hydroxides** of metals
- When **alkalis** are added to water, they form negative hydroxide ions (OH^-)
- The presence of the OH^- ions is what makes the aqueous solution an alkali

Example: Sodium Hydroxide

$$NaOH\,(s) \rightarrow Na^+\,(aq) + OH^-\,(aq)$$

Typical reactions of bases

Bases and acids

- When bases react with an acid, a neutralisation reaction occurs
- Acids and bases react together in a neutralisation reaction and produce a salt and water:

Acid + Base → Salt + Water

Examples of reaction between bases and acids:

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride and water	$Mg(OH)_2 + 2HCl \longrightarrow MgCl_2 + 2H_2O$
Sulfuric acid	Magnesium sulfate and water	$MgO + H_2SO_4 \longrightarrow MgSO_4 + H_2O$
Nitric acid	Magnesium nitrate and water	$Mg(OH)_2 + 2HNO_3 \longrightarrow Mg(NO_3)_2 + H_2O$

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Alkalis and ammonium salts

- Ammonium salts undergo decomposition when warmed with an alkali
- Even though ammonia is itself a weak base, it is very **volatile** and can easily be displaced from the salt by another alkali
- A salt, water and ammonia are produced

Example:

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$$NH_4Cl + NaOH \rightarrow NaCl + H_2O + NH_3$$

- This reaction is used as a chemical test to confirm the presence of the ammonium ion (NH_4^+)
- Alkali is added to the substance with gentle warming followed by the test for ammonia gas using damp red litmus paper
- The damp litmus paper will turn from **red** to **blue** if ammonia is present



Your notes

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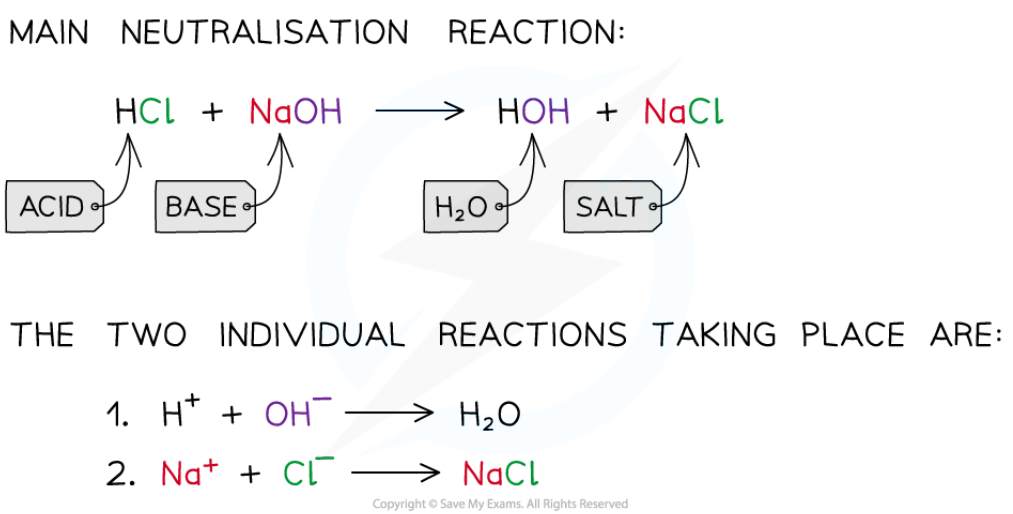
The Ions in Acids & Alkalis

Your notes

Neutralisation Reactions

- When **acids** are added to water, they form positively charged **hydrogen ions** (H^+)
- The presence of H^+ ions is what makes a solution acidic
- When **alkalis** are added to water, they form negative **hydroxide ions** (OH^-)
- The presence of the OH^- ions is what makes the aqueous solution an alkali
- The pH scale is a numerical scale which is used to show how **acidic** or **alkaline** a solution is, in other words it is a measure of the amount of ions present in the solution

- A neutralisation reaction occurs when an acid reacts with an alkali
- When these substances react together in a neutralisation reaction, the **H^+** ions react with the **OH^-** ions to produce water
- For example, when hydrochloric acid is neutralised, sodium chloride and water are produced:



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- The net ionic equation of **acid-alkali neutralisations**, and what leads to a neutral solution, since water has a pH of 7, is:

$$H^+ + OH^- \rightarrow H_2O$$

Exam Tip

Not all reactions of acids are neutralisations. For example, when a metal reacts with an acid, although a salt is produced there is no water formed so it does not fit the definition of neutralisation.

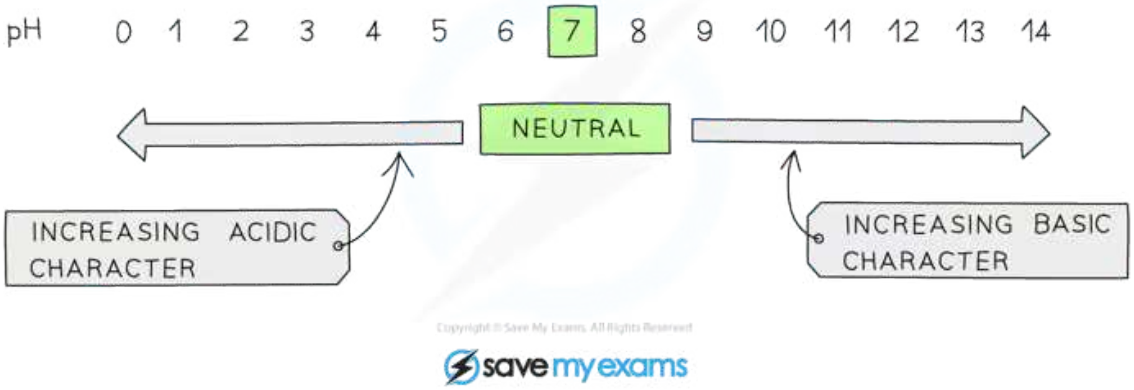
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Hydrogen Ion Concentration & pH

The pH scale

Your notes

- The pH scale is a numerical scale which is used to show how **acidic** or **alkaline** a solution is
- It goes from 1 - 14 (extremely acidic substances can have values of below 1)
- All acids have pH values of **below** 7, all alkalis have pH values of **above** 7
- The **lower** the pH then the **more acidic** the solution is
- The **higher** the pH then the **more alkaline** the solution is
- A solution with a pH of 7, such as water, is described as being **neutral**



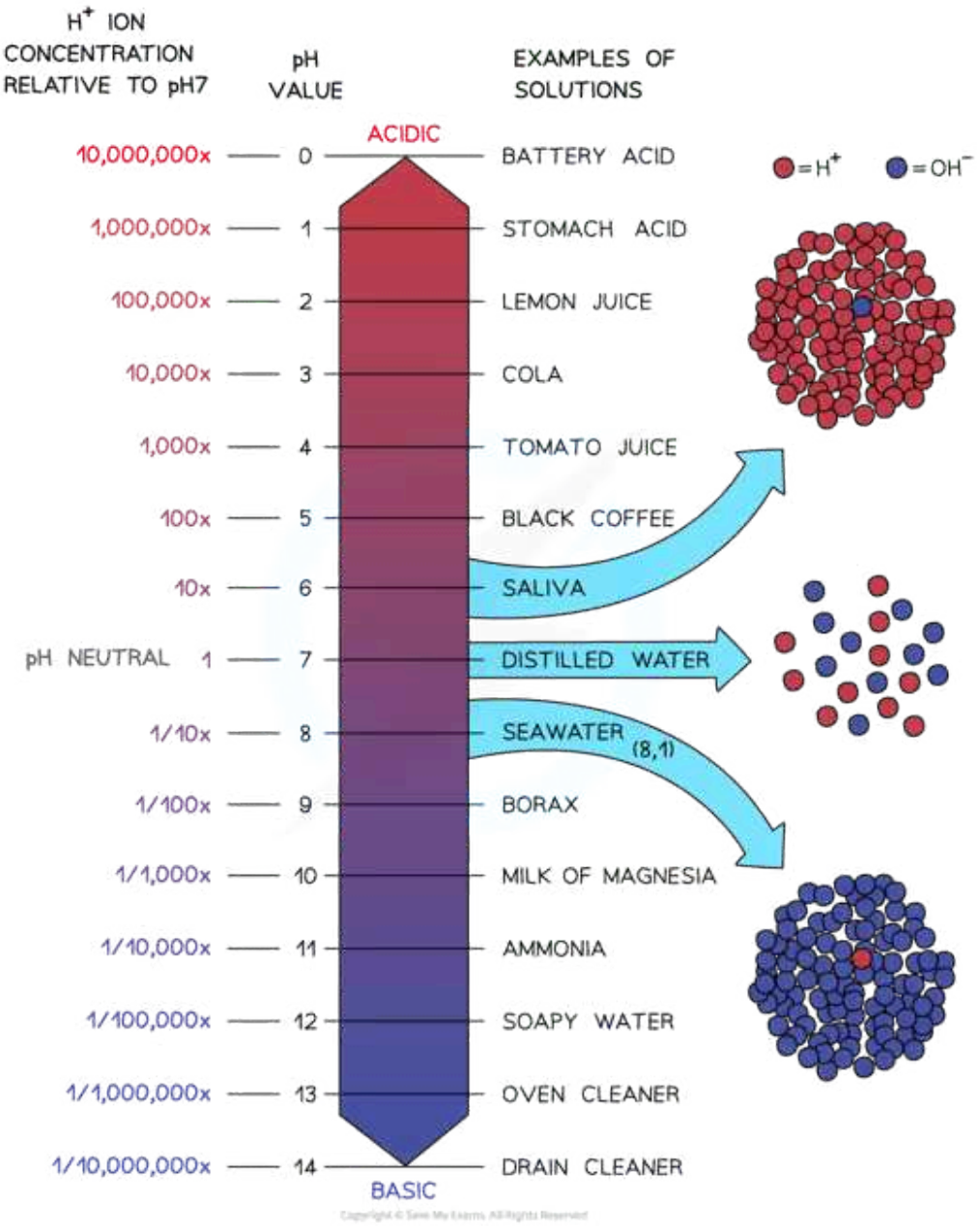
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The pH scale showing acidity, neutrality and alkalinity

- We have already seen that acids are substances that contain hydrogen ions in solution
- The more hydrogen ions the stronger the acid, but the lower the pH
- The higher the concentration of hydroxide ions in a solution the higher the pH
- So pH is a measure of the concentration of H^+ ions in solution, but they have an inverse relationship

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- The pH scale is **logarithmic**, meaning that each change of 1 on the scale represents a change in concentration by a **factor** of **10**
 - Therefore an acid with a pH of 3 has ten times the concentration of H^+ ions than an acid of pH 4
 - An acid with a pH of 2 has 10 x 10 = 100 times the concentration of H^+ ions than an acid with a pH of 4

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Exam Tip

Acid strength is reflected in how many hydrogen ions are in solution. The more hydrogen ions the lower the pH and vice-versa.

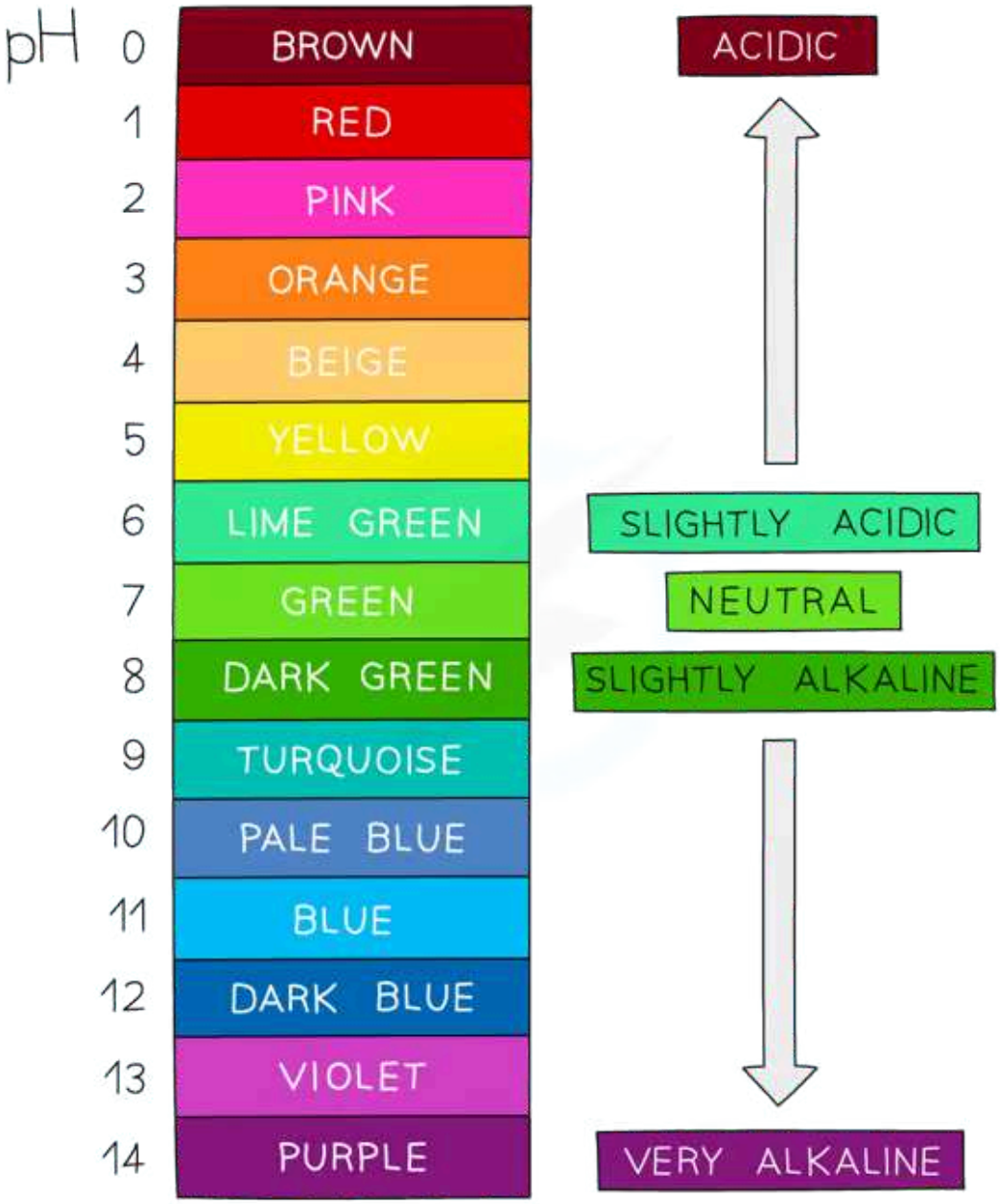
Your notes

Universal indicator

- Universal indicator is a mixture of different **indicators** which is used to measure the pH
- A drop is added to the solution and the colour is matched with a colour chart which indicates the pH which matches specific colours

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The pH scale with the Universal Indicator colours which can be used to determine the pH of a solution



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Proton Transfer, Strong & Weak Acids

Your notes

Proton Transfer, Strong & Weak Acids

Proton transfer

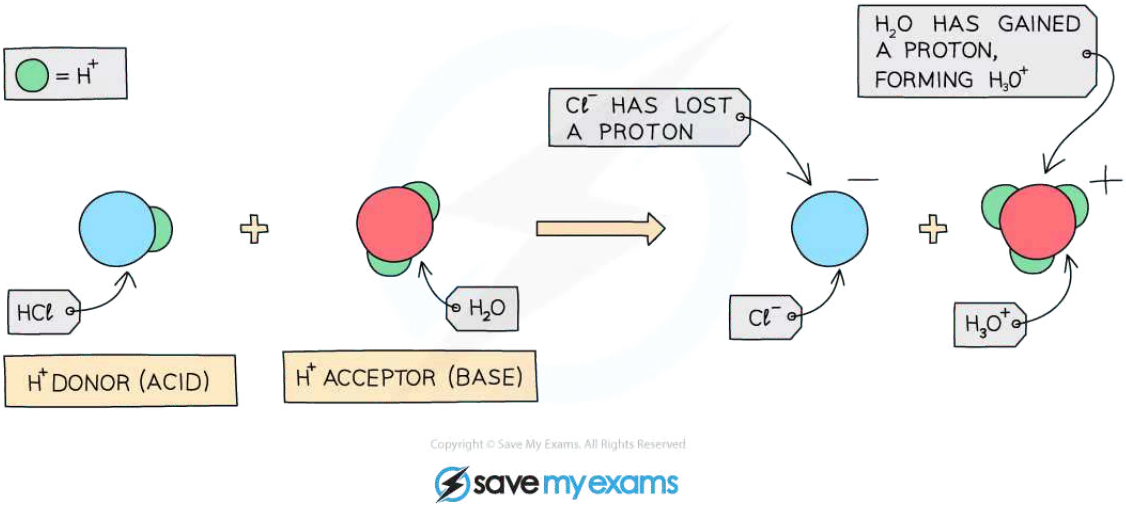
- The earlier definition of an acid and a base can be extended
- In terms of proton transfer, we can further define each substance in how they interact with protons

Acids

- Acids are **proton donors** as they ionise in solution producing protons, which are H^+ ions
- These H^+ ions make the aqueous solution acidic

Bases

- Bases are **proton acceptors** as they accept the protons which are donated by the acid



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Diagram showing the role of acids and bases in the transfer of protons – here water acts as a base as it accepts a proton

Strong acids

- Acids can be either **strong** or **weak**, depending on how many H^+ ions they produce when dissolved in water
- Strong acids **completely dissociate** (or ionise) in water, producing a **high** amount of H^+ ions and producing solutions with a **very low pH**

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- Strong acids include HCl and H_2SO_4
- Example of a strong acid: hydrochloric acid

$$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$$

Your notes

Weak acids

- Weak acids **partially dissociate** (or ionise) in water, producing **fewer H^+ ions** in solution that strong acids and producing solutions with pH values which are closer to the **middle** of the pH scale, whilst still being below 7
- Weak acids include organic acids such as ethanoic acid, CH_3COOH
- For weak acids, there is usually an equilibrium set-up between the molecules and their ions once they have been added to water
- Example of a weak acid: propanoic acid

$$CH_3CH_2COOH \rightleftharpoons H^+ + CH_3CH_2COO^-$$

- The equilibrium lies to the **left**, indicating a high concentration of intact acid molecules, with a low concentration of H^+ ions in the solution

Effect of concentration on strong and weak acids

- A concentrated solution of an acid is one that contains a **higher number** of acid **molecules** per dm^3 of solution
- It does not necessarily mean that the acid is strong though, as it may be made from a weak acid which does not dissociate completely
- For example a dilute solution of HCl will be more acidic than a concentrated solution of ethanoic acid, since most of the HCl molecules dissociate but very few of the CH_3COOH do

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Cambridge O Level Chemistry

Your notes

7.2 Oxides

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Classifying Oxides

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Classifying Oxides


Your notes

Classifying Oxides

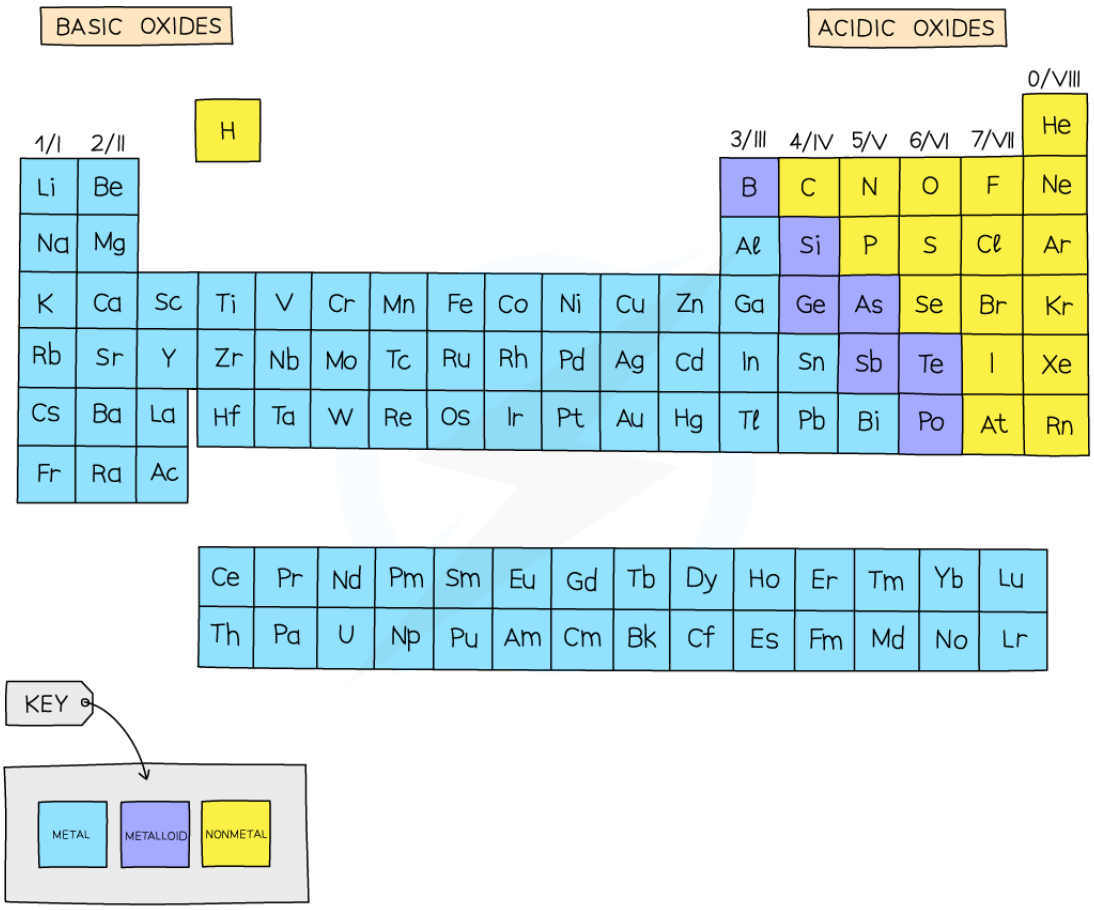
- Oxides are compounds made from one or more atoms of oxygen combined with one other element
- Examples of oxides include: MgO, ZnO, K_2O, CO_2, SO_2, H_2O
- Oxides can be classified based on their acid-base characteristics

Acidic and basic oxides

- Acidic and basic oxides have different properties and values of pH
- The difference in their pH stems from whether they are bonded to a **metal** or a **non-metal** element
- The metallic character of the element influences the acidic or basic behaviour of the molecule

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Metals form basic oxides while non-metals form acidic oxides

Acidic oxides

- Acidic oxides are formed when a **non-metal** element combines with oxygen
- They react with bases to form a salt and water
- When dissolved in water they produce an **acidic** solution with a **low** pH
- Common examples include CO_2, SO_2, NO_2 and SiO_2

Basic oxides

- Basic oxides are formed when a **metal** element combines with oxygen
- They react with acids to form a salt and water
- When dissolved in water they produce a **basic** solution with a **high** pH

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- Common examples include CuO and CaO


Your notes

Amphoteric Oxides

Amphoteric oxides

- Amphoteric oxides are a curious group of oxides that can behave as **both acidic** and **basic**, depending on whether the other reactant is an acid or a base
- In both cases, a salt and water are formed
- Two of the most common amphoteric oxides are **zinc oxide**, **ZnO** and **aluminum oxide**, **Al_2O_3**
- The **hydroxides** of both of these elements also behave amphoterically
- Example of aluminium oxide behaving as a base:

$$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$$

- Example of aluminium oxide behaving as an acid:

$$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$$

- This acidic and basic behaviour is not easily explained by donating or accepting protons. A separate theory called the Lewis acid-base theory can identify acids or bases in these situations, but is not required for this course

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Cambridge O Level Chemistry

Your notes

7.3 Preparation of Salts

Contents

- Preparing Soluble Salts
- Preparing Insoluble Salts
- Solubility Rules
- Hydrated & Anhydrous Salts

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Preparing Soluble Salts

Your notes

Preparing Soluble Salts

Salts

- A salt is a compound that is formed when the **hydrogen atom** in an acid is replaced by a **metal**
- For example if we replace the H in HCl with a potassium atom, then the salt potassium chloride is formed, KCl
- Salts are an important branch of chemistry due to the varied and important uses of this class of compounds
- These uses include fertilisers, batteries, cleaning products, healthcare products and fungicides

Naming salts

- The name of a salt has two parts
 - The first part comes from the **metal**, **metal** oxide or **metal** carbonate used in the reaction
 - The second part comes from the **acid**
- The name of the salt can be determined by looking at the reactants
- For example hydrochloric acid always produces salts that end in chloride and contain the **chloride** ion, Cl^-
-
 - **Sodium** hydroxide reacts with hydro**chloric** acid to produce sodium chloride
 - **Zinc** oxide reacts with **sulf**uric acid to produce **zinc sulfate**Other examples:

Preparing salts

- Some salts can be extracted by mining but others need to be prepared in the laboratory
- The method used depends on the solubility of the salt being prepared

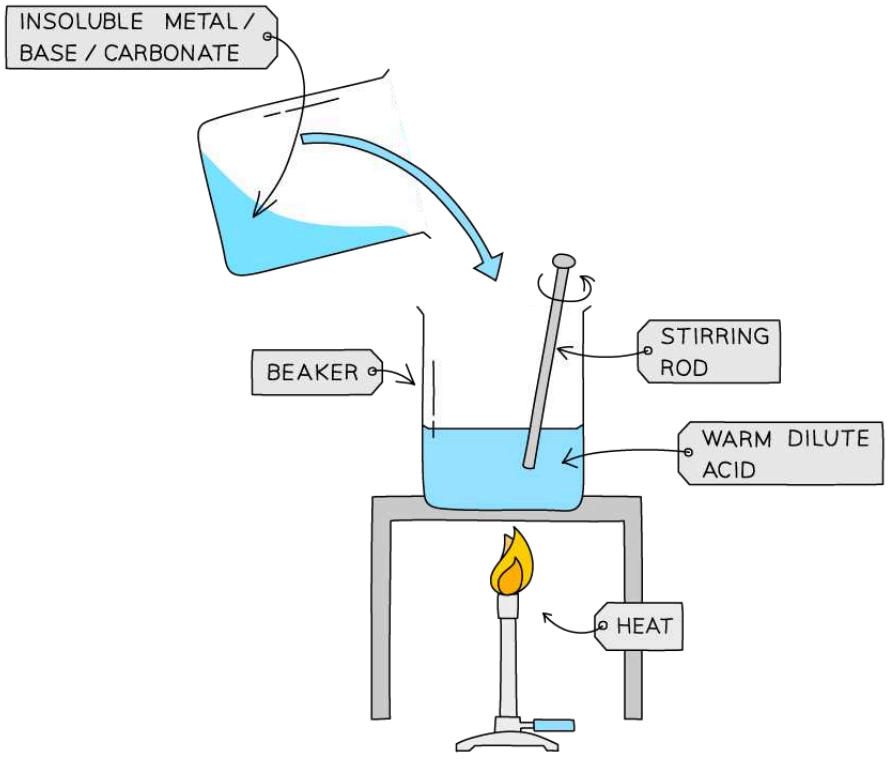
Preparing soluble salts

Method A: Adding acid to a solid metal, insoluble base or insoluble carbonate

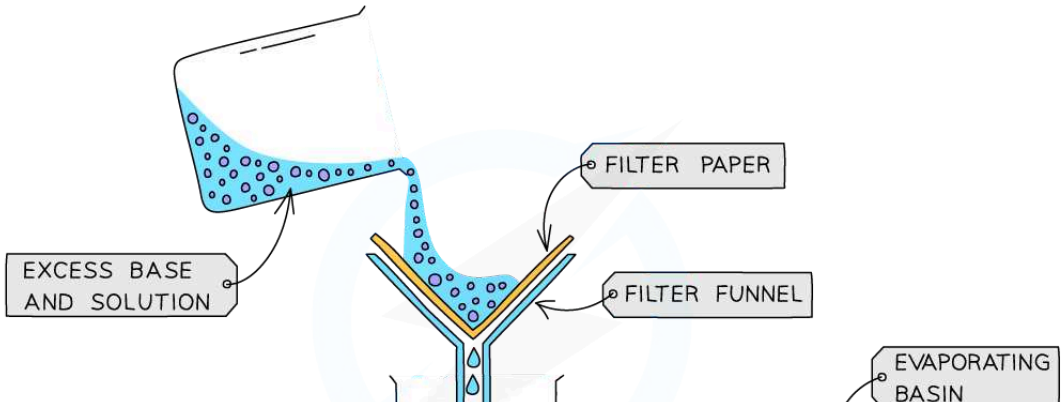
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Your notes

1 HEAT ACID UNTIL WARM, THEN ADD METAL / BASE / CARBONATE, STIRRING CONSTANTLY UNTIL IT STOPS DISAPPEARING



2 FILTER MIXTURE TO REMOVE EXCESS BASE, TRANSFER SOLUTION TO EVAPORATING BASIN.



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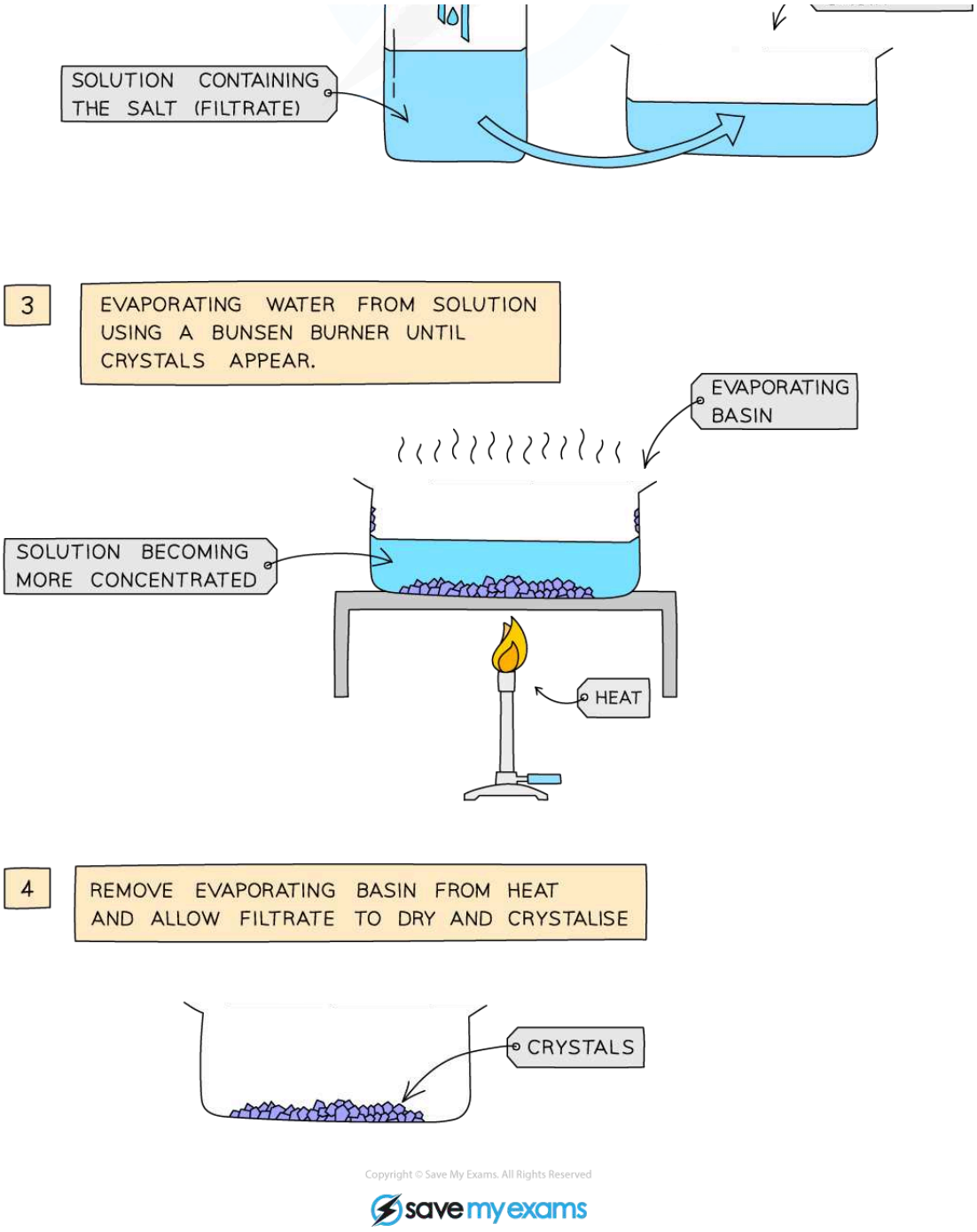


Diagram showing the preparation of soluble salts

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Your notes

Method:

- Add dilute acid into a beaker and heat using a bunsen burner flame
- Add the insoluble metal, base or carbonate, a little at a time, to the warm dilute acid and stir until the base is in excess (i.e. until the base stops disappearing and a suspension of the base forms in the acid)
- Filter the mixture into an evaporating basin to remove the excess base
- Heat the solution to evaporate water and to make the solution saturated. Check the solution is saturated by dipping a cold, glass rod into the solution and seeing if crystals form on the end
- Leave the filtrate in a warm place to dry and crystallize
- Decant excess solution and allow crystals to dry or blot to dry with filter paper

Example: Preparation of pure, hydrated copper(II) sulfate crystals using method A

Acid = dilute sulfuric acid

Insoluble base = copper(II) oxide

Method:

- Add dilute sulfuric acid into a beaker and heat using a bunsen burner flame
- Add copper(II) oxide (insoluble base), a little at a time to the warm dilute sulfuric acid and stir until the copper (II) oxide is in excess (stops disappearing)
- Filter the mixture into an evaporating basin to remove the excess copper(II) oxide
- Leave the filtrate in a warm place to dry and crystallize
- Decant excess solution
- Blot crystals dry with filter paper

Equation of reaction:

$$\text{copper(II) oxide} + \text{sulfuric acid} \rightarrow \text{copper(II) sulphate} + \text{water}$$

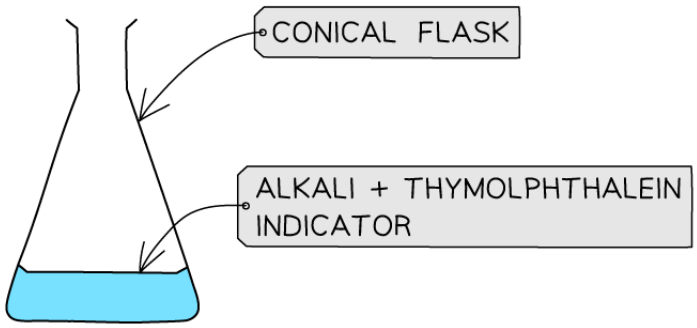
$$CuO\,(s) + H_2SO_4\,(aq) \rightarrow CuSO_4\,(aq) + H_2O\,(l)$$

Method B: Reacting a dilute acid and alkali (soluble base)

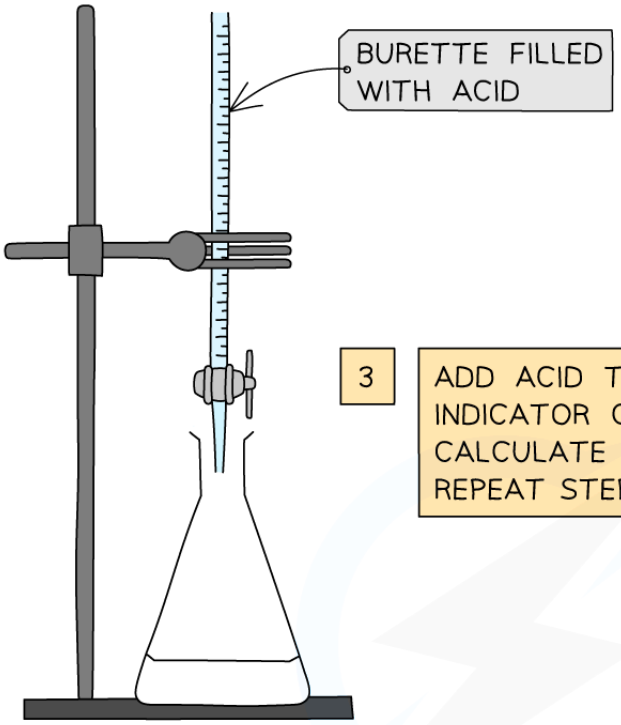
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Your notes

1 ADD ALKALI + INDICATOR TO CONICAL FLASK USING A PIPETTE



2 ADD ACID TO BURETTE, NOTING THE STARTING VOLUME



3 ADD ACID TO ALKALI SLOWLY UNTIL INDICATOR CHANGES COLOUR.
CALCULATE VOLUME OF ACID ADDED.
REPEAT STEPS 1–3 WITHOUT INDICATOR

4 TRANSFER SOLUTION TO AN EVAPORATING BASIN, HEAT TO PARTIALLY EVAPORATE WATER

EVAPORATING

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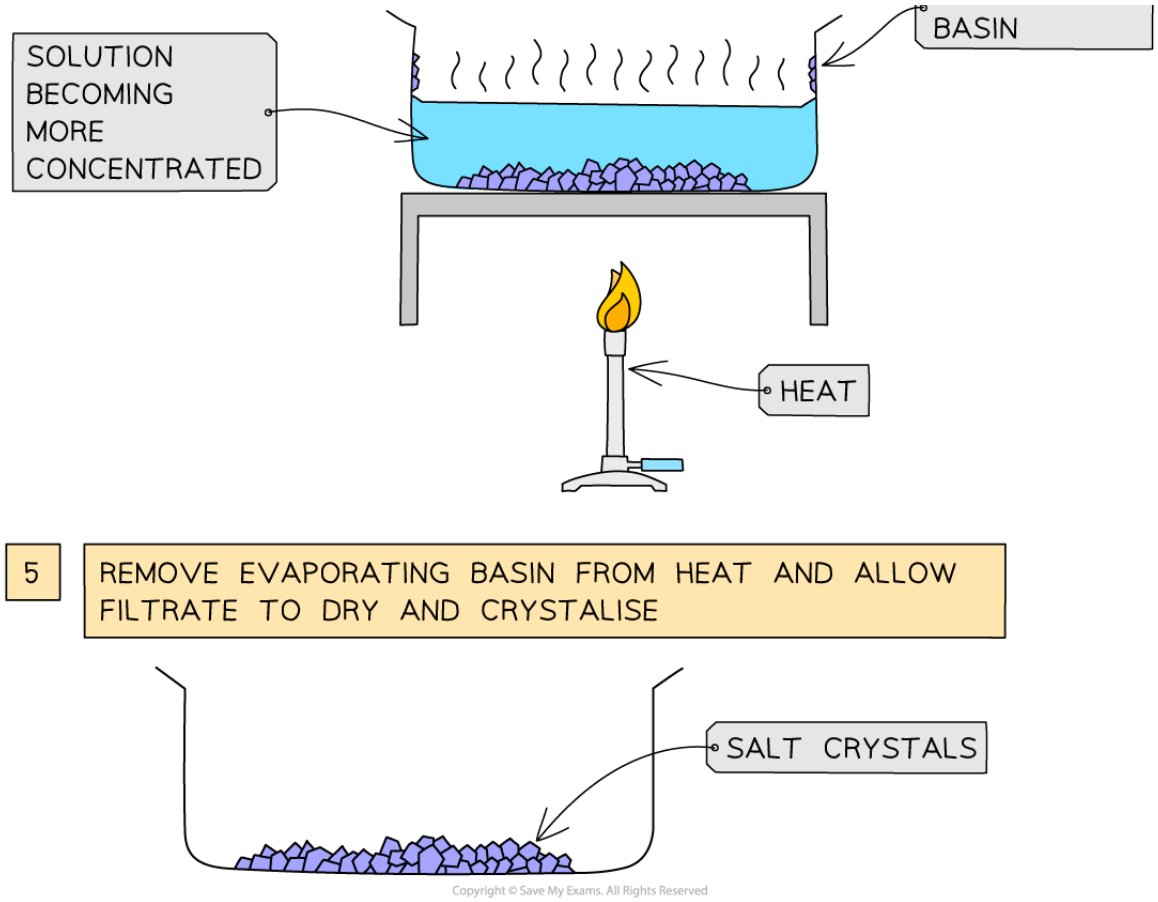


Diagram showing the apparatus needed to prepare a salt by titration

Method:

- Use a pipette to measure the alkali into a conical flask and add a few drops of indicator (thymolphthalein or methyl orange)
- Add the acid into the burette and note the starting volume
- Add the acid very slowly from the burette to the conical flask until the indicator changes to the appropriate colour
- Note and record the final volume of acid in the burette and calculate the volume of acid added (starting volume of acid - final volume of acid)
- Add this same volume of acid into the same volume of alkali without the indicator
- Heat the resulting solution in an evaporating basin to partially evaporate, leaving a saturated solution (crystals just forming on the sides of the basin or on a glass rod dipped in and then removed)
- Leave to crystallise, decant excess solution and allow crystals to dry

Your notes

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Preparing Insoluble Salts


Your notes

Preparing Insoluble Salts

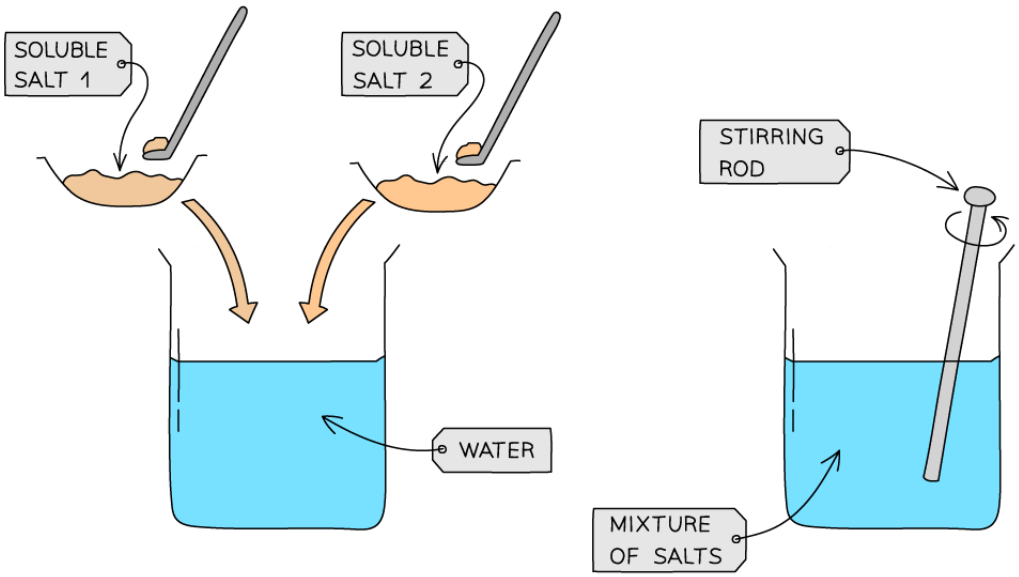
- Insoluble salts can be prepared using a precipitation reaction
- The solid salt obtained is the precipitate, thus in order to successfully use this method the solid salt being formed must be insoluble in water, and the reactants must be soluble

Using two soluble reactants

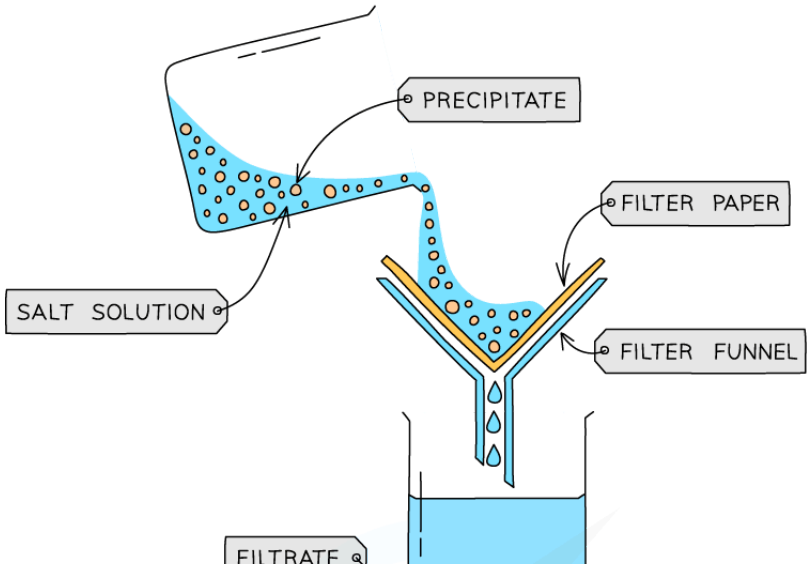
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Your notes

1 ADD SOLUBLE SALTS TO WATER AND MIX

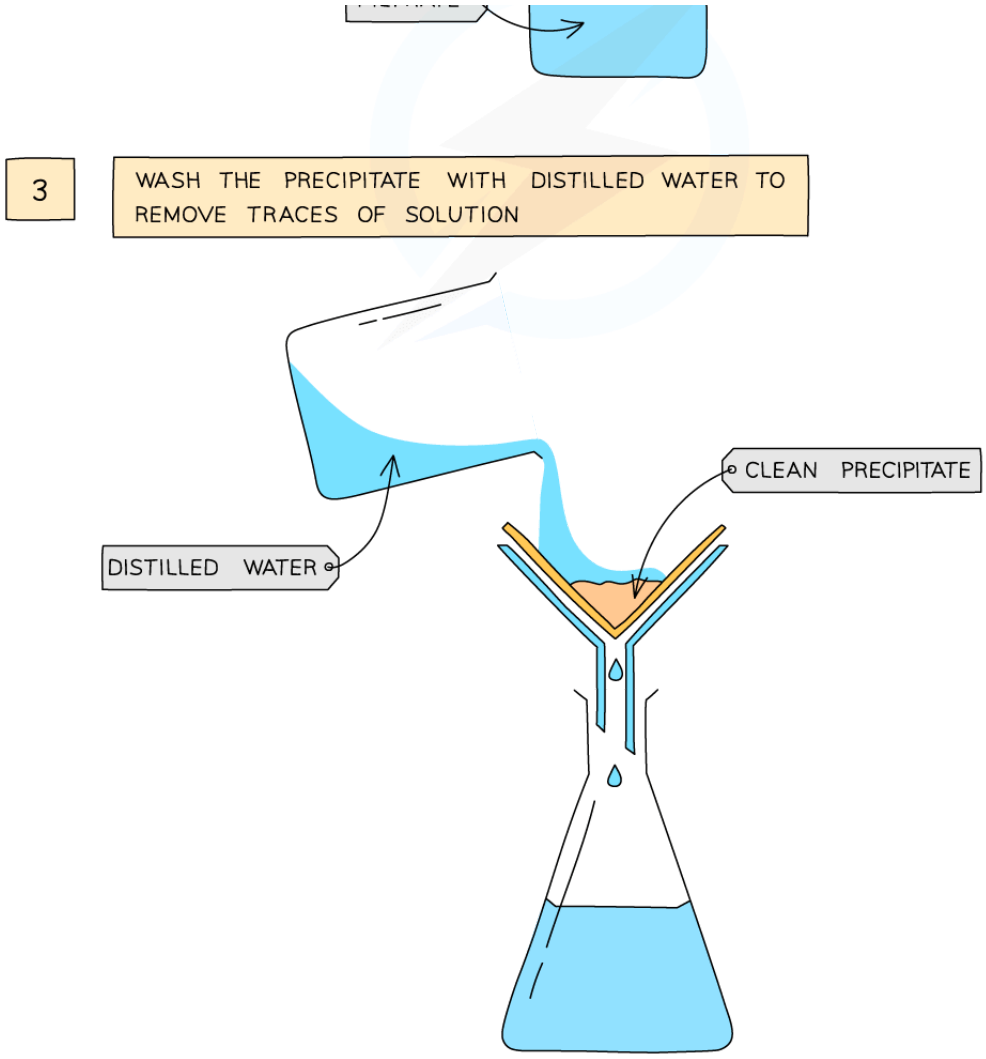


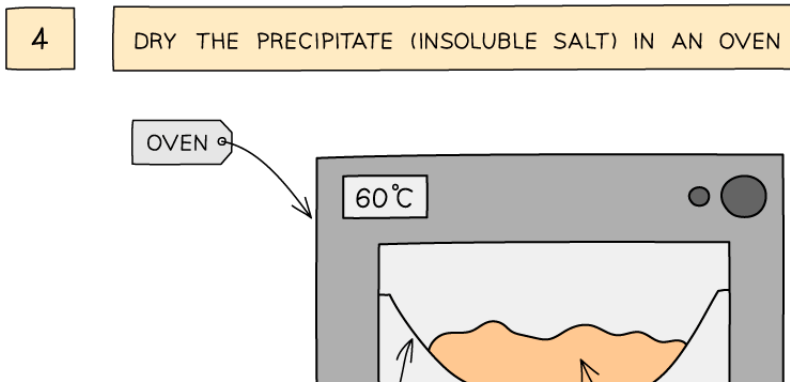
2 FILTER TO REMOVE PRECIPITATE FROM THE MIXTURE



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Diagram showing the filtration of the mixture to remove the precipitate

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Method:

- Dissolve soluble salts in water and mix together using a stirring rod in a beaker
- Filter to remove precipitate from mixture
- Wash filtrate with distilled water to remove traces of other solutions
- Leave in an oven to dry

Example: Preparation of pure, dry lead(II) sulfate crystals using a precipitation reaction

Soluble Salt 1 = lead(II) nitrate

Soluble Salt 2 = potassium sulfate

Method:

- Dissolve lead(II) nitrate and potassium sulfate in water and mix together using a stirring rod in a beaker
- Filter to remove precipitate from mixture
- Wash precipitate with distilled water to remove traces of potassium nitrate solution
- Leave in an oven to dry

Equation of reaction:

lead(II) nitrate + potassium sulfate → lead(II) sulfate + potassium nitrate

$Pb(NO_3)_2\,(aq) + K_2SO_4\,(aq) \rightarrow PbSO_4\,(s) + 2KNO_3\,(aq)$

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Solubility Rules


Your notes

Solubility Rules

- Salts are prepared by different methods, depending on whether the salt is soluble or insoluble so it is important to know the solubility of salts

Solubility of the common salts

Salts	Soluble	Insoluble
Sodium, potassium and ammonium	All	None
Nitrates	All	None
Chlorides	Most are soluble	Silver and lead(II)
Sulfates	Most are soluble	Barium, calcium and lead(II)
Carbonates	Carbonates of sodium, potassium and ammonium	Most are insoluble
Hydroxides	Hydroxides of sodium, potassium, ammonium and calcium (calcium hydroxide is partially soluble	Most are insoluble

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Hydrated & Anhydrous Salts


Your notes

Hydrated & Anhydrous Salts

- When salts are being prepared, some water can be retained within the structure of the salt during the crystallisation process
- This affects the crystal's shape and colour
- Salts that contain water within their structure are called hydrated salts
- Anhydrous salts are those that contain no water in their structure
- A common example is copper(II) sulfate which crystallises forming the salt hydrated copper(II) sulfate, which is blue
- When it is heated, the water from its structure is removed, forming anhydrous copper(II) sulfate, which is white
- The hydrated salt has been **dehydrated** to form the anhydrous salt
- This reaction can be reversed by adding water to anhydrous copper(II) sulfate:

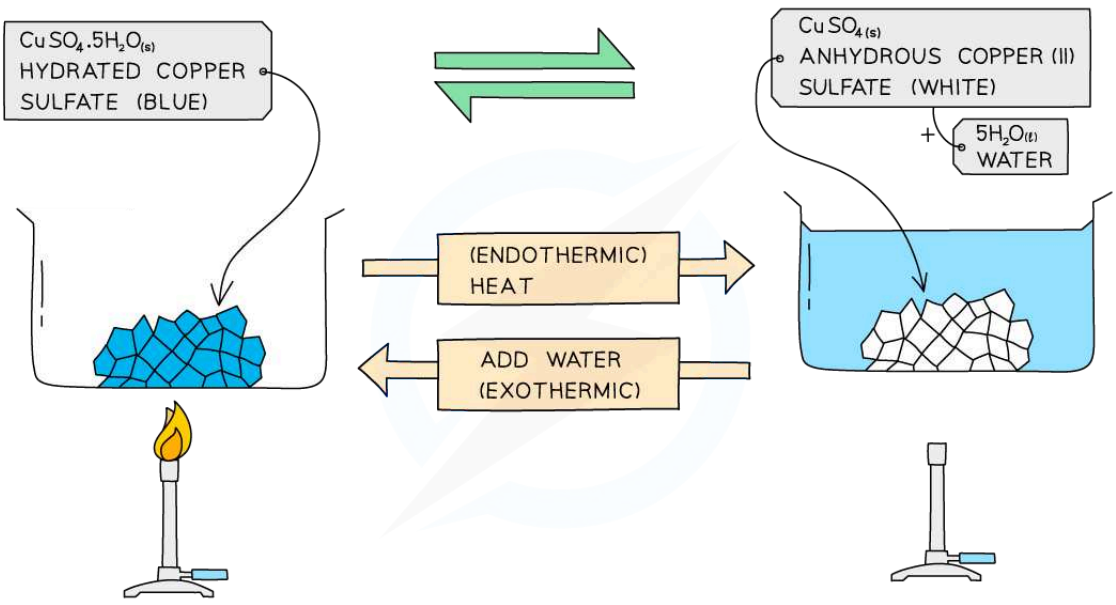
hydrated copper(II) sulfate ⇌ anhydrous copper(II) sulfate + water

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Water of Crystallisation

Your notes

- Water molecules included in the structure of some salts during the crystallisation process are known as **water of crystallisation**
- A compound that contains water of crystallisation is called a **hydrated compound**
- When writing the chemical formula of hydrated compounds, the water of crystallisation is separated from the main formula by a **dot**
 - E.g. hydrated copper(II) sulfate is $CuSO_4 \cdot 5H_2O$
 - Hydrated cobalt(II) chloride is $CoCl_2 \cdot 6H_2O$
- The formula shows the number of moles of water contained within one mole of the hydrated salt
 - E.g. hydrated copper(II) sulfate, $CuSO_4 \cdot 5H_2O$, contains 5 moles of water in 1 mole of hydrated salt
- A compound which doesn't contain water of crystallisation is called an **anhydrous compound**
 - E.g. anhydrous copper(II) sulfate is $CuSO_4$
 - Anhydrous cobalt(II) chloride is $CoCl_2$
- The conversion of anhydrous compounds to hydrated compounds is **reversible** by heating the hydrated salt:
 - Anhydrous to hydrated salt:
 - **$CuSO_4 + 5H_2O \rightarrow CuSO_4 \cdot 5H_2O$**
 - Hydrated to anhydrous salt (by heating):
 - **$CuSO_4 \cdot 5H_2O \rightarrow CuSO_4 + 5H_2O$**



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Diagram showing the dehydration of hydrated copper II) sulfate

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Cambridge O Level Chemistry

Your notes

8.1 Arrangement of Elements

Contents

- The Periodic Table
- Periodic Trends

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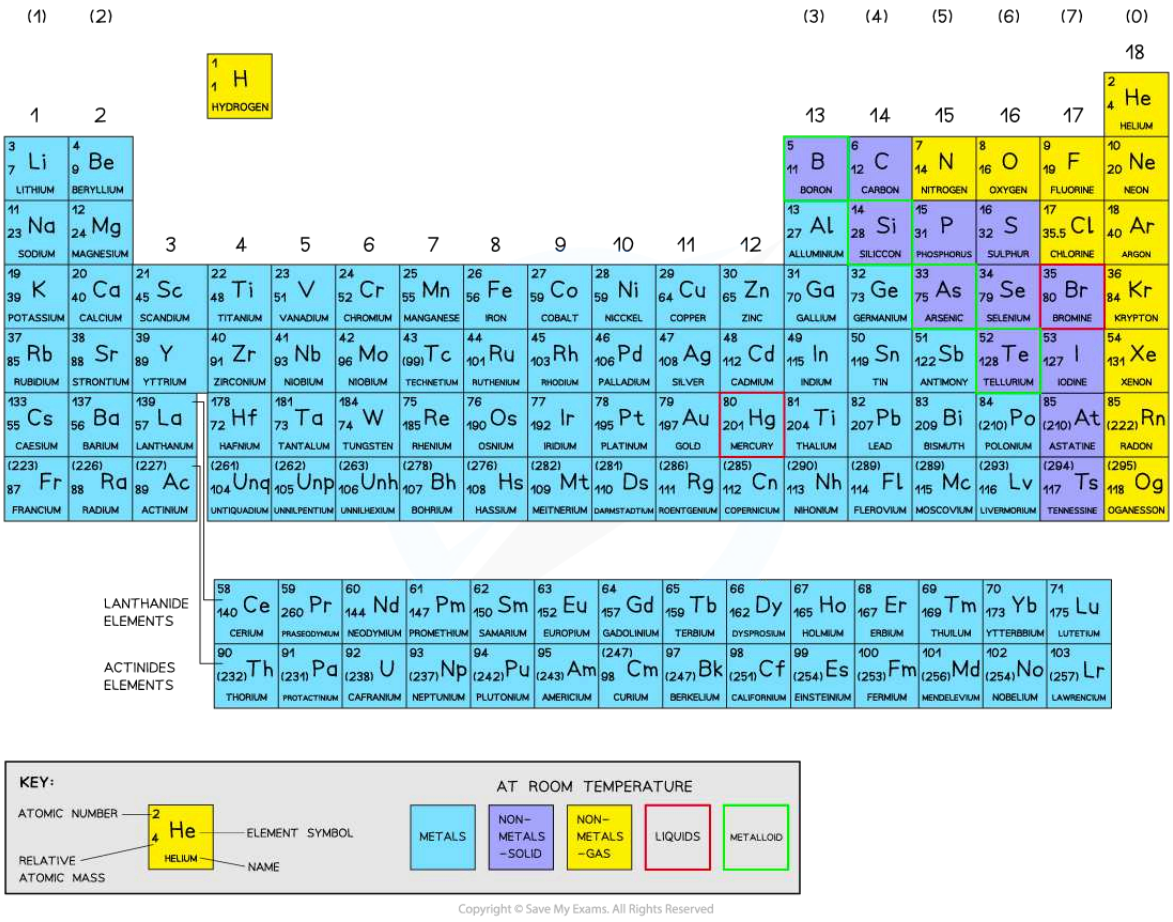
The Periodic Table


Your notes

The Periodic Table

- There are over 100 chemical elements which have been isolated and identified
 - Each element has one proton **more** than the element preceding it
 - This is done so that elements end up in columns with other elements which have similar properties
 - Elements are arranged on the periodic table in order of **increasing atomic number**
 - The table is arranged in vertical columns called **groups** and in rows called **periods**
- **Period:** These are the horizontal rows that show the number of shells of electrons an atom has and are numbered from 1 - 7
 - E.g. elements in period 2 have two electron shells, elements in period 3 have three electron shells
- **Group:** These are the vertical columns that show how many outer electrons (also known as valency electrons) each atom has and are numbered from I – VII, with a final group called Group 0 (instead of Group VIII)
 - E.g. Group IV elements have atoms with 4 electrons in the outermost shell, Group VI elements have atoms with 6 electrons in the outermost shell and so on
- The group number can help determine the charge that metal and non-metal ions form
- For metals, the group number corresponds to the number of electrons it will lose to achieve a full outer shell and the charge of the metal ion
 - E.g. sodium is in **Group I,** it will lose 1 electron and form an ion with a **1+ charge**
 - Magnesium is in **Group** II, it will lose 2 electrons and form an ion with a **2+ charge**
- For non-metals in Group VII and VI, they will gain 1 and 2 electrons respectively to gain a full outer shell
 - E.g. non-metals in **Group VII** gain 1 electron to form ions with a **1- charge**
 - Non-metals in **Group VI** gain 2 electrons to form ions with a **2- charge**

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All elements are arranged in the order of increasing atomic number from left to right

Valency

- Valency (or combining power) tells you how many bonds an atom can make with another atom or how many electrons its atoms lose, gain or share, to form a compound
 - E.g. carbon has a valancy of 4 as it is in Group IV so a single carbon atom can share 4 electrons to make 4 single bonds or 2 double bonds
- The following valencies apply to elements in each group:

Your notes

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GROUP	VALENCY
I	1
II	2
III	3
IV	4
V	3
VI	2
VII	1
VIII	0

Your notes

Exam Tip

An easier way of remembering which number is the mass number and which is the atomic is:

Mass Number = The **mass**ive number i.e the larger of the two numbers.

The atomic number must be the smaller number.

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Periodic Trends

Your notes

The Metallic Character of Elements

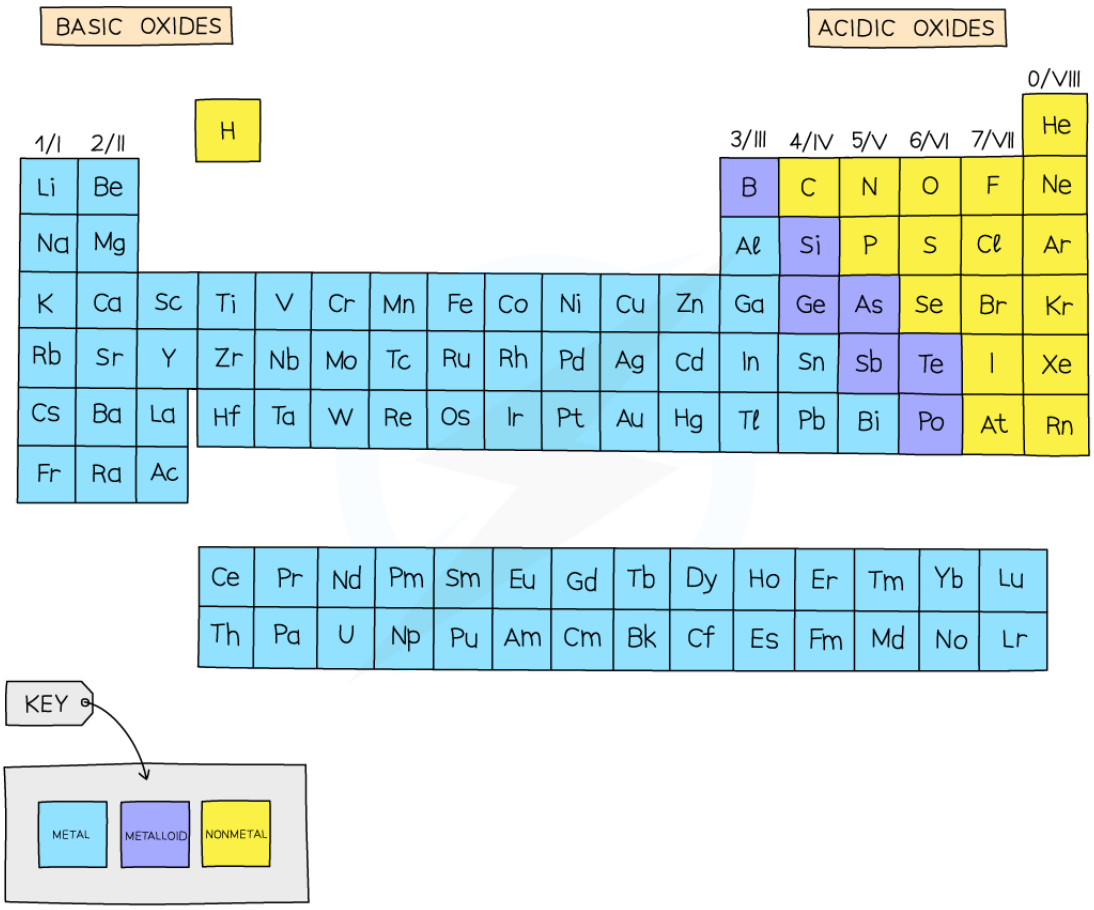
- The metallic character of the elements **decreases** as you move across a Period on the Periodic Table, from **left to right**, and it **increases** as you move down a Group
- This trend occurs due to atoms more **readily accepting** electrons to fill their valence shells rather than losing them to have the previous, already full, electron shell as their outer shell
- Metals occur on the **left-hand** side of the Periodic Table and non-metals on the **right-hand** side
- Between the metals and the non-metals lie the elements which display some properties of **both**
- These elements are referred to as metalloids or semi-metals

Properties of metals and non-metals

PROPERTY	METALS	NON-METALS
ELECTRON ARRANGEMENT	1 – 3 (MORE IN PERIODS 5 & 6) OUTER SHELL ELECTRONS	4 – 7 ELECTRONS IN THE OUTER SHELL
BONDING	METALLIC DUE TO LOSS OF OUTER SHELL ELECTRONS	COVALENT BY SHARING OF OUTER SHELL ELECTRONS
ELECTRICAL CONDUCTIVITY	GOOD CONDUCTORS OF ELECTRICITY	POOR CONDUCTORS OF ELECTRICITY
TYPE OF OXIDE	BASIC OXIDES (A FEW ARE AMPHOTERIC)	ACIDIC OXIDES (SOME ARE NEUTRAL)
REACTION WITH ACIDS	MANY REACT WITH ACIDS	DO NOT REACT WITH ACIDS
PHYSICAL CHARACTERISTICS	MALLEABLE, CAN BE BENT AND SHAPED HIGH MELTING AND BOILING POINT	FLAKY, BRITTLE LOW MELTING AND BOILING POINT

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A zig-zag line in this diagram separates the metals on the left, from the non-metals on the right

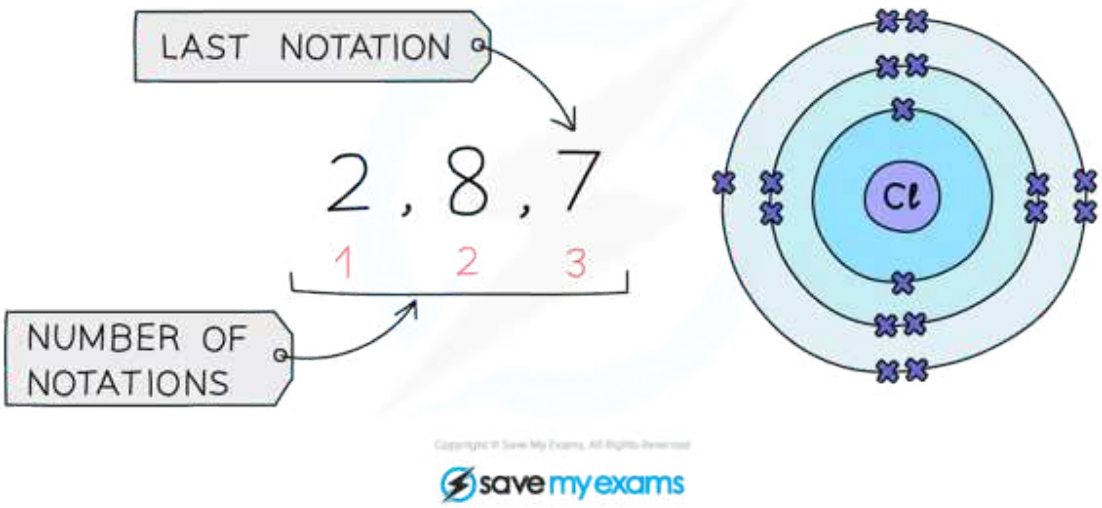
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Periodic Trends & Electronic Configuration

Your notes

- The electronic configuration is the arrangement of electrons into shells for an atom (e.g: the electronic configuration of carbon is 2,4)
- There is a link between the electronic configuration of the elements and their position on the Periodic Table
- The number of notations in the electronic configuration will show the number of occupied shells of electrons the atom has, showing the **period**
- The last notation shows the number of outer electrons the atom has, showing the **group** number

Example: Electronic configuration of **chlorine**:



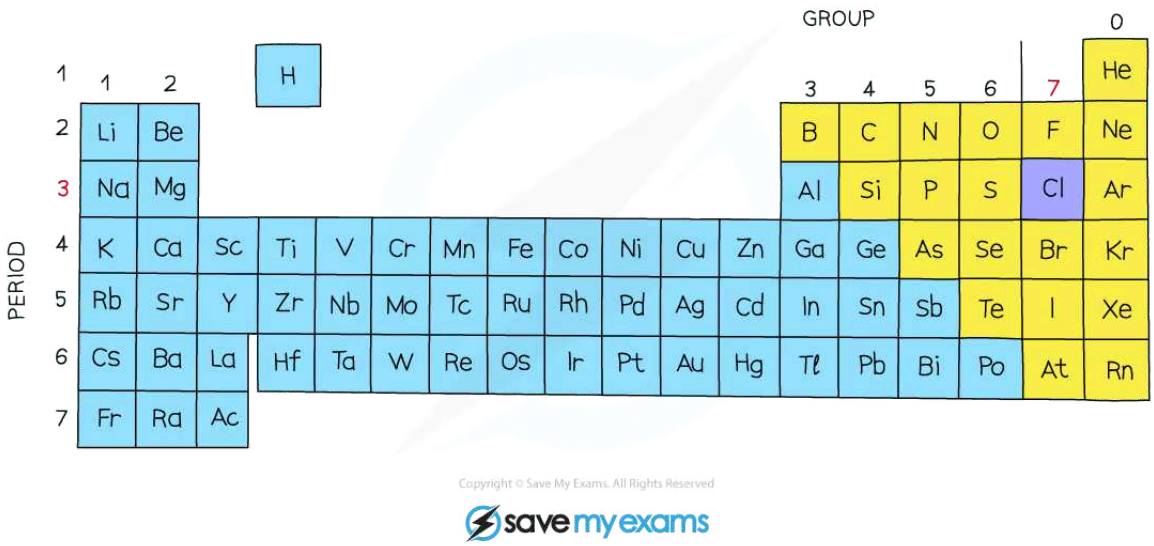
The electronic configuration of chlorine as it should be written

Period: The red numbers at the bottom show the number of notations which is 3, showing that a chlorine atom has 3 shells of electrons.

Group: The final notation, which is 7 in the example, shows that a chlorine atom has 7 outer electrons and is in Group VII

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Your notes



The position of chlorine on the Periodic Table

- Elements in the same group in the Periodic Table have **similar chemical properties**
- When atoms collide and react, it is the **outermost electrons** that interact
- The similarity in their chemical properties stems from having the **same number** of electrons in their outer shell
- For example, both lithium and sodium are in Group I and can react with elements in Group VII to form an ionic compound (charges of Group I ions are 1+, charges of Group VII ions are 1–) by reacting in a similar manner and each donating one electron to the Group VII element
- As you look down a group, a full shell of electrons is added to each subsequent element
 - Lithium's electronic configuration: 2,1
 - Sodium's electronic configuration: 2,8,1
 - Potassium's electronic configuration: 2,8,8,1

Exam Tip

Electronic configurations can be shown with the numbers separated by commas or by full stops. In this course commas are used, but you will often see full stops used elsewhere. Both are accepted.

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Predicting Properties


Your notes

- Because there are **patterns** in the way the elements are arranged on the Periodic Table, there are also **patterns** and **trends** in the chemical behaviour of the elements and their physical properties
- These trends in properties occur down groups and across the periods of the Periodic Table
- As a result, we can use the Periodic Table to predict properties such as:
 - boiling point
 - melting point
 - density
 - reactivity
- Some common properties / trends in properties include:
 - Group I elements react very quickly with water
 - Noble gases are unreactive
 - Transition elements are denser than Group I elements
 - Reactivity decreases going down Group VII
 - Melting point decreases going down Group I
- In this way the Periodic Table can be used to **predict** how a particular element will behave

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Identifying Trends

Your notes

- Using given information about elements, we can identify trends in properties
- An example of when this might be used is to determine the trend in reactivity of Group I metals
- The table below shows the reactions of the first three elements in Group I with water

Observations of Lithium, Sodium, and Potassium with Water

Element	Reaction
Lithium	$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$ ∘ Reaction slower than with sodium ∘ Bubbles of hydrogen gas ∘ Lithium's melting point is higher and heat isn't produced so quickly, so the lithium doesn't melt
Sodium	$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$ ∘ Bubbles of hydrogen gas ∘ Melts into a shiny ball that dashes around the surface ∘ Floats on water because it is less dense ∘ Melts because sodium has a low melting point and a lot of heat is made in the reaction ∘ Hydrogen is evolved which causes the ball of sodium to move around the surface of the water ∘ White trail of sodium hydroxide produced which dissolves in the water producing a highly alkaline solution
Potassium	$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$ ∘ Reacts more violently than sodium ∘ Bubbles of gas ∘ Melts into a shiny ball that dashes around the surface ∘ Enough heat produced so burns with a lilac–coloured flame

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- The observations show that reactivity of the Group I metals increases as you go down the group
- Using this information we can predict the trend going further down Group I for the elements rubidium, caesium and francium
- As the reactivity of alkali metals increases down the group, rubidium, caesium and francium will react more vigorously with air and water than lithium, sodium and potassium

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- Lithium will be the **least** reactive metal in the group at the top, and francium will be the **most** reactive at the bottom
- Francium is rare and radioactive so is difficult to confirm predictions

Table to Show the Predicted Reaction of other Group I Elements with Water

ELEMENT	REACTION
RUBIDIUM	– EXPLODES WITH SPARKS – RUBIDIUM HYDROXIDE PRODUCED.
CAESIUM	– VIOLENT EXPLOSION DUE TO RAPID PRODUCTION OF HEAT AND HYDROGEN – CAESIUM HYDROXIDE PRODUCED.
FRANCIUM	– TOO REACTIVE TO PREDICT

Your notes

Exam Tip

You may be asked to identify other trends in chemical or physical properties of Group I metals, given appropriate data.

Firstly, ensure that the metals and associated data are written in either descending or ascending order according the their position in the Group. Then look for general patterns in the data.

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Cambridge O Level Chemistry

Your notes

8.2 Group I Properties

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Group I Properties

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Group I Properties

Your notes

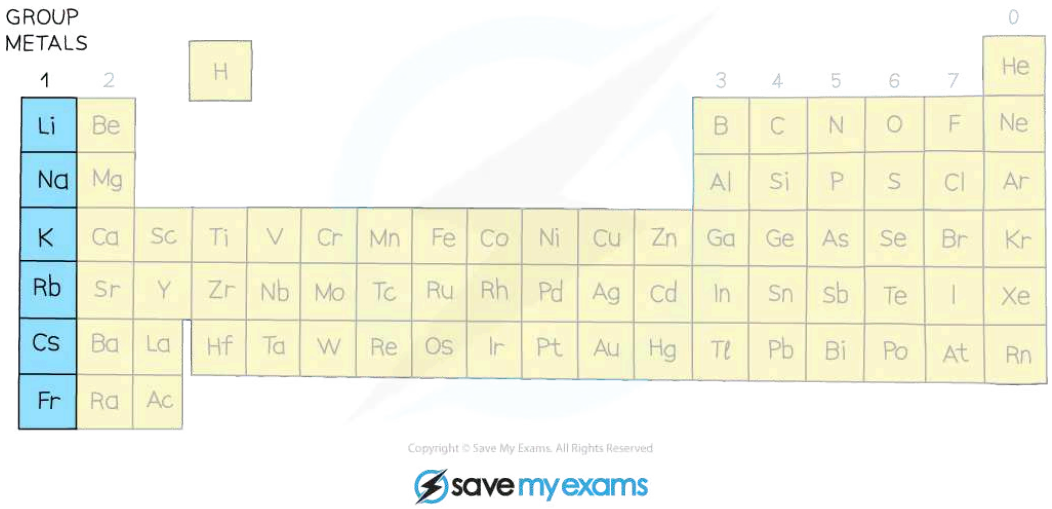
Group I Properties & Trends: Basics

The Group I metals

- The Group I metals are also called the alkali metals as they form **alkaline solutions** with high pH values when reacted with water
- Group I metals are lithium, sodium, potassium, rubidium, caesium and francium
- They all contain just **one electron** in their outer shell

Physical properties of the Group I metals

- The Group I metals:
 - Are soft and easy to cut, getting even **softer** and **denser** as you move down the Group (sodium and potassium do not follow the trend in density)
 - Have **shiny** silvery surfaces when freshly cut
 - Conduct **heat** and **electricity**
 - They all have **low** melting points and **low** densities compared to other metals, and the melting point **decreases** as you move down the Group; some would melt on a hot day



The alkali metals lie on the far left-hand side of the Periodic Table

Chemical properties of the Group I metals

- They react readily with oxygen and water vapour in air so they are stored under **oil** to stop them from reacting

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- Group I metals will react similarly with water, reacting vigorously to produce an **alkaline** metal hydroxide solution and **hydrogen** gas
- The Group I metals get more reactive as you look down the group, so only the first three metals are allowed in schools for demonstrations

Your notes

Reactions of the Group I metals and water

Element	Reaction
Lithium	$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$ ∘ Reaction slower than with sodium ∘ Bubbles of hydrogen gas ∘ Lithium's melting point is higher and heat isn't produced so quickly, so the lithium doesn't melt
Sodium	$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$ ∘ Bubbles of hydrogen gas ∘ Melts into a shiny ball that dashes around the surface ∘ Floats on water because it is less dense ∘ Melts because sodium has a low melting point and a lot of heat is made in the reaction ∘ Hydrogen is evolved which causes the ball of sodium to move around the surface of the water ∘ White trail of sodium hydroxide produced which dissolves in the water producing a highly alkaline solution
Potassium	$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$ ∘ Reacts more violently than sodium ∘ Bubbles of gas ∘ Melts into a shiny ball that dashes around the surface ∘ Enough heat produced so burns with a lilac-coloured flame

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Predicting the Properties of Group I Elements

Your notes

- Knowing the reactions of elements at the top of the group allows you to predict the properties of other elements further down Group I

Properties of other Alkali Metals (Rubidium, Caesium and Francium)

- As the reactivity of alkali metals increases down the group, rubidium, caesium and francium will react more vigorously with air and water than lithium, sodium and potassium
- Lithium will be the **least** reactive metal in the group at the top, and francium will be the **most** reactive at the bottom
- Francium is rare and radioactive so is difficult to confirm predictions
- For example the reactions with water can be predicted:

Predicting the Reaction with Water

ELEMENT	REACTION
RUBIDIUM	– EXPLODES WITH SPARKS – RUBIDIUM HYDROXIDE PRODUCED.
CAESIUM	– VIOLENT EXPLOSION DUE TO RAPID PRODUCTION OF HEAT AND HYDROGEN – CAESIUM HYDROXIDE PRODUCED.
FRANCIUM	– TOO REACTIVE TO PREDICT

- You can also look at other properties such as boiling point, melting point and density of Group I elements and use them to predict whether the other properties are likely to be larger or smaller going down the group

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Your notes

8.3 Group VII Properties

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- Group VII Properties
- Group VII Displacement Reactions

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Group VII Properties

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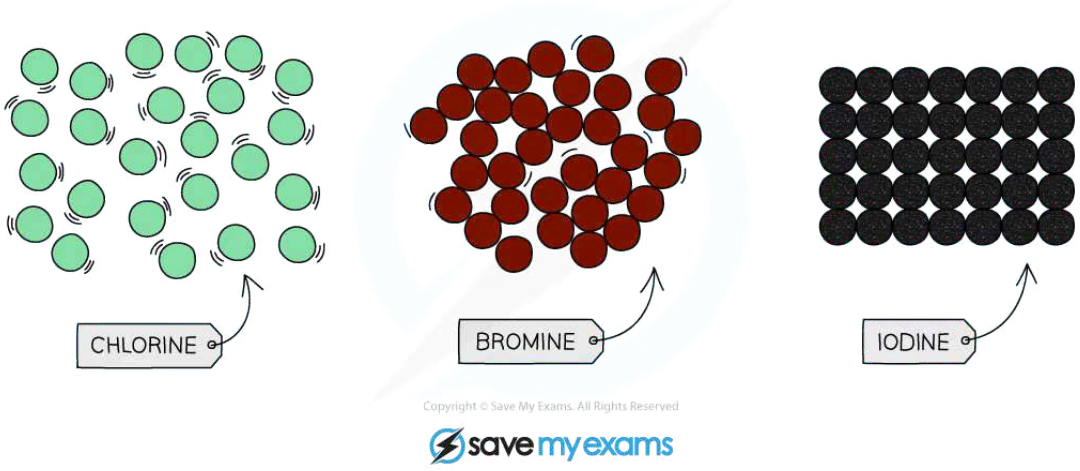
Group VII Properties & Trends

The halogens

- These are the Group VII non-metals that are **poisonous** and include fluorine, chlorine, bromine, iodine and astatine
- Halogens are **diatomic**, meaning they form molecules of **two** atoms
 - The formulae of the halogens are F_2, Cl_2, Br_2, I_2 and At_2
- All halogens have seven electrons in their outer shell
- They form halide ions by gaining one more electron to complete their outer shells
- Fluorine is not allowed in schools so observations and experiments tend to only involve chlorine, bromine and iodine

Properties of the halogens

- At room temperature (20 °C), the physical state of the halogens changes as you go down the group
- Chlorine is a **pale yellow-green gas**, bromine is a **red-brown liquid** and iodine is a **grey-black solid**
- This demonstrates that the **density** of the halogens **increases** as you go **down the group:**



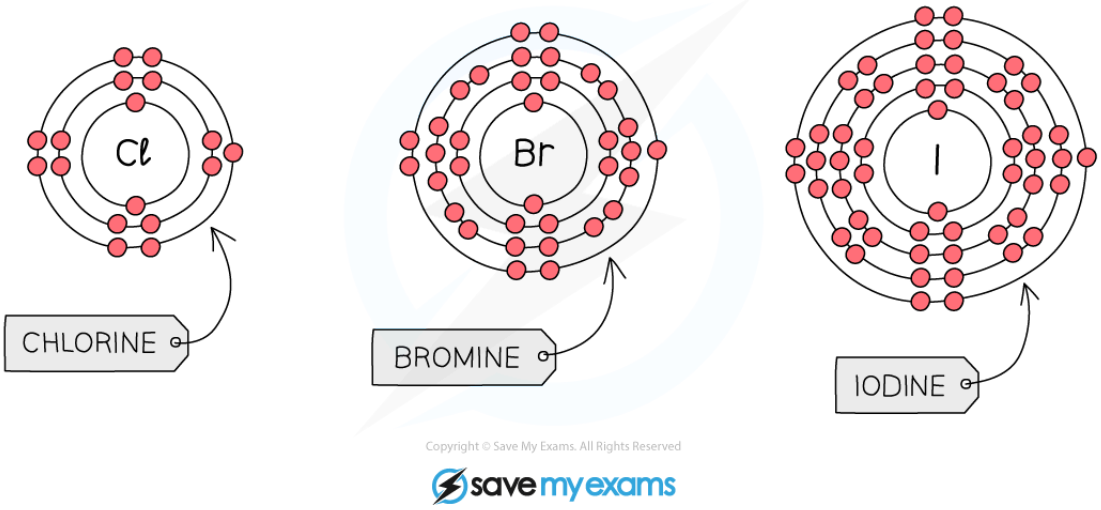
The physical state of the halogens at room temperature

- **Reactivity** of Group VII non-metals **increases** as you go **up** the group (this is the opposite trend to that of Group I)
- Each outer shell contains seven electrons and when the halogen reacts, it will need to gain one outer electron to get a full outer shell of electrons
- As you go up Group VII, the number of shells of electrons **decreases** (period number decreases moving up the Periodic Table)

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- This means that the outer electrons are **closer** to the nucleus so there are **stronger** electrostatic forces of attraction, which help to attract the extra electron needed
- This allows an electron to be attracted more readily, so the **higher** up the element is in Group VII then the **more reactive** it is

Your notes



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Diagram showing the electronic configuration of the first three elements in Group VII

Exam Tip

Solid iodine, iodine in solution and iodine vapour are different colours. Solid iodine is dark **grey-black**, iodine vapour is **purple** and aqueous iodine is **brown**.

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Predicting Group VII Properties

- You may be given information about some elements and asked to predict the properties of other elements in the group
- The information you might be given could be in relation to melting/boiling point or physical state/density so it is useful to know the trends in properties going down the group

Melting and boiling point

- The melting and boiling point of the halogens **increases** as you go down the group
- Fluorine is at the top of Group VII so will have the **lowest** melting and boiling point
- Astatine is at the bottom of Group VII so will have the **highest** melting and boiling point

Physical states

- The halogens become **denser** as you go down the group
- Fluorine is at the top of Group VII so will be a **gas**
- Astatine is at the bottom of Group VII so will be a **solid**

Colour

- The colour of the halogens becomes **darker** as you go down the group
- Fluorine is at the top of Group VII so the colour will be **lighter**, so fluorine is **yellow**
- Astatine is at the bottom of Group VII so the colour will be **darker**, so astatine is **black**

Exam Tip

You can be asked to identify trends in chemical or physical properties of the Group VII elements, given appropriate data.

Firstly, make sure that you have placed the elements and associated data in either ascending or descending order according to their position in Group VII. Then look for any general patterns in the data.


Your notes

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Group VII Displacement Reactions


Your notes

Group VII Displacement Reactions

- A halogen displacement reaction occurs when a more reactive halogen displaces a less reactive halogen from an aqueous solution of its halide
- The reactivity of Group VII non-metals increases as you move up the group
- Out of the three commonly used halogens, chlorine, bromine and iodine, chlorine is the most reactive and iodine is the least reactive

Colour of Halogens in Aqueous Solutions

AQUEOUS SOLUTION	COLOUR
CHLORINE	VERY PALE GREEN, BUT USUALLY APPEARS COLOURLESS AS IT IS VERY DILUTE
BROMINE	ORANGE BUT WILL TURN YELLOW WHEN DILUTED
IODINE	BROWN

Halogen displacement reactions

Chlorine and bromine

- If you add chlorine solution to colourless potassium bromide solution, the solution becomes orange as bromine is formed
- Chlorine is **above** bromine in Group VII so is more reactive
- Chlorine will therefore **displace** bromine from an aqueous solution of the metal bromide
- The least reactive halogen always ends up in the elemental form

potassium bromide + chlorine → potassium chloride + bromine

$$2KBr\,(aq) + Cl_2\,(aq) \rightarrow 2KCl\,(aq) + Br_2\,(aq)$$

Bromine and iodine

- Bromine is above iodine in Group VII so is **more** reactive
- Bromine will therefore displace iodine from an aqueous solution of metal iodide
- The solution will turn brown as iodine is formed

magnesium iodide + bromine → magnesium bromide + iodine

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$MgI_2(aq) + Br_2(aq) \rightarrow MgBr_2(aq) + I_2(aq)$

Your notes

Summary table of displacement reactions

	Chlorine (Cl_2)	Bromine (Br_2)	Iodine (I_2)
Potassium chloride (KCl)	X	No reaction	No reaction
Potassium bromide (KBr)	Chlorine displaces the bromide ions. Yellow–orange colour of bromine is seen	X	No reaction
Potassium iodide (KI)	Chlorine displaces the iodide ions. Brown colour of iodine is seen	Bromine displaces the iodide ions:. Brown colour of iodine is seen	X

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Exam Tip

Iodine solid, solution and vapour are different colours. Solid iodine is dark **grey-black**, iodine vapour is **purple** and aqueous iodine is **brown**.

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Cambridge O Level Chemistry

Your notes

8.4 Transition Elements

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Transition Elements

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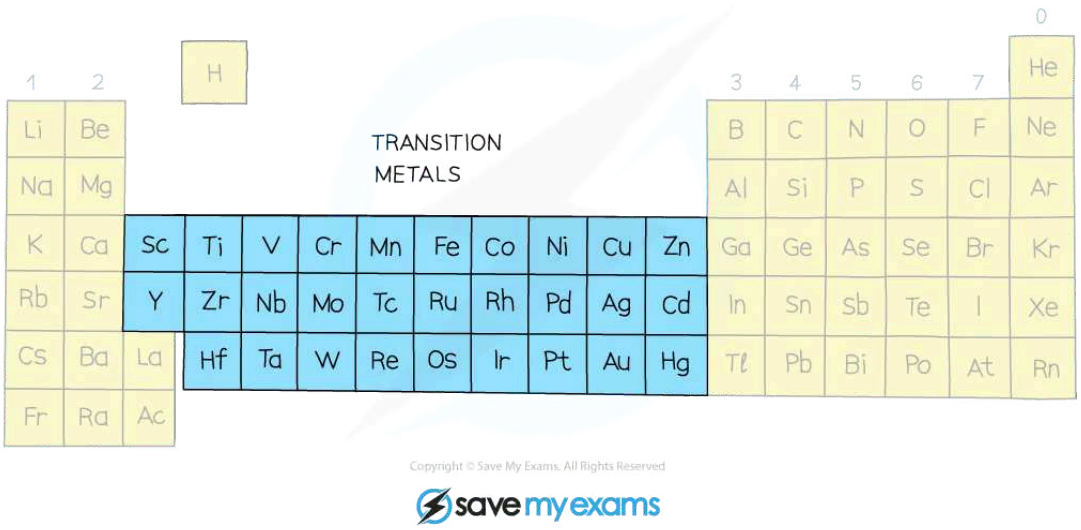
Transition Elements

Your notes

Transition Elements

General properties of the transition elements

- They are very **hard** and strong metals and are good conductors of **heat** and **electricity**
- They have very **high melting** points and are highly **dense** metals
- For example, the melting point of titanium is 1,688°C whereas potassium in Group I melts at only 63.5°C, slightly warmer than the average cup of hot chocolate!
- The transition elements form **coloured** compounds and often have more than one oxidation state, such as iron readily forming compounds of both Fe^{2+} and Fe^{3+}
- These coloured compounds are responsible for the pigments in many paints and the colours of gemstones and rocks
- Transition elements, as elements or in compounds, are often used as **catalysts** to improve the rate or reaction in industrial processes
 - Transition element catalysts of platinum or rhodium are also used in car exhausts in the 'catalytic convertor' to reduce the levels of nitrous oxides and carbon monoxide produced



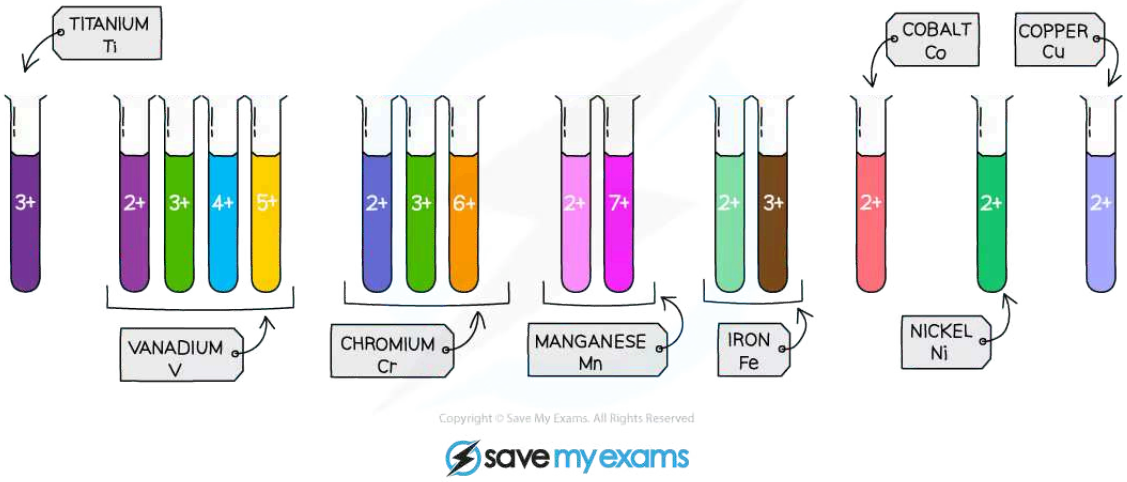
The transition elements on the Periodic Table

- The transition elements have more than one **oxidation number**, as they can lose a different number of electrons, depending on the chemical environment they are in
- For example. iron either:
 - Lose **two** electrons to form Fe^{2+} so has an oxidation number of +2
 - Loses **three** electrons to form Fe^{3+} so has an oxidation number of +3

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- Compounds containing transition elements in different oxidation states will have different **properties** and colours

Your notes



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Ions of the same element can have different oxidation numbers forming different colours

Uses of the transition elements

- The transition elements are used extensively as **catalysts** due to their ability to interchange between a range of oxidation states
- This allows them to form complexes with reagents which can easily **donate** and **accept** electrons from other chemical species within a reaction system
- They are used in **medicine** and **surgical** applications such as limb and joint replacement (titanium is often used for this as it can bond with bones due to its high biocompatibility)
- They are also used to form coloured compounds in **dyes** and **paints**, **stained** glass **jewellery**

Exam Tip

Although scandium and zinc are in the transition element area of the Periodic Table, they are not considered transition elements as they do **not** form coloured compounds and have only **one** oxidation state.

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8.5 Noble Gases

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Noble Gases

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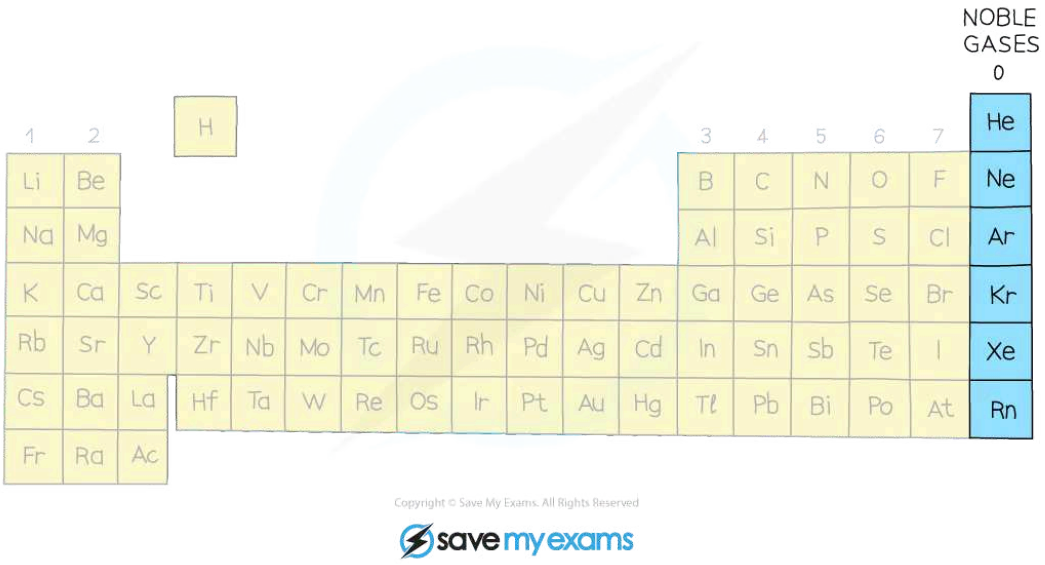
Noble Gases

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Noble Gases Properties & Electronic Configuration

The Noble Gases

- The noble gases are in Group VIII (or Group 0); they are non-metals and have very **low melting** and **boiling** points
- They are all **monoatomic**, **colourless** gases
- The Group 0 elements all have **full outer shells**
- This electronic configuration is **extremely stable** so these elements are unreactive and are inert
- Electronic configurations of the noble gases:
 - He: 2
 - Ne: 2,8
 - Ar: 2,8,8
 - Kr: 2,8,18,8
 - Xe: 2,8,18,18,8



Noble gases are inert (unreactive) as they have a full outer shell of electrons so do not easily lose or gain electrons

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9.1 Properties of Metals

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Properties of Metals

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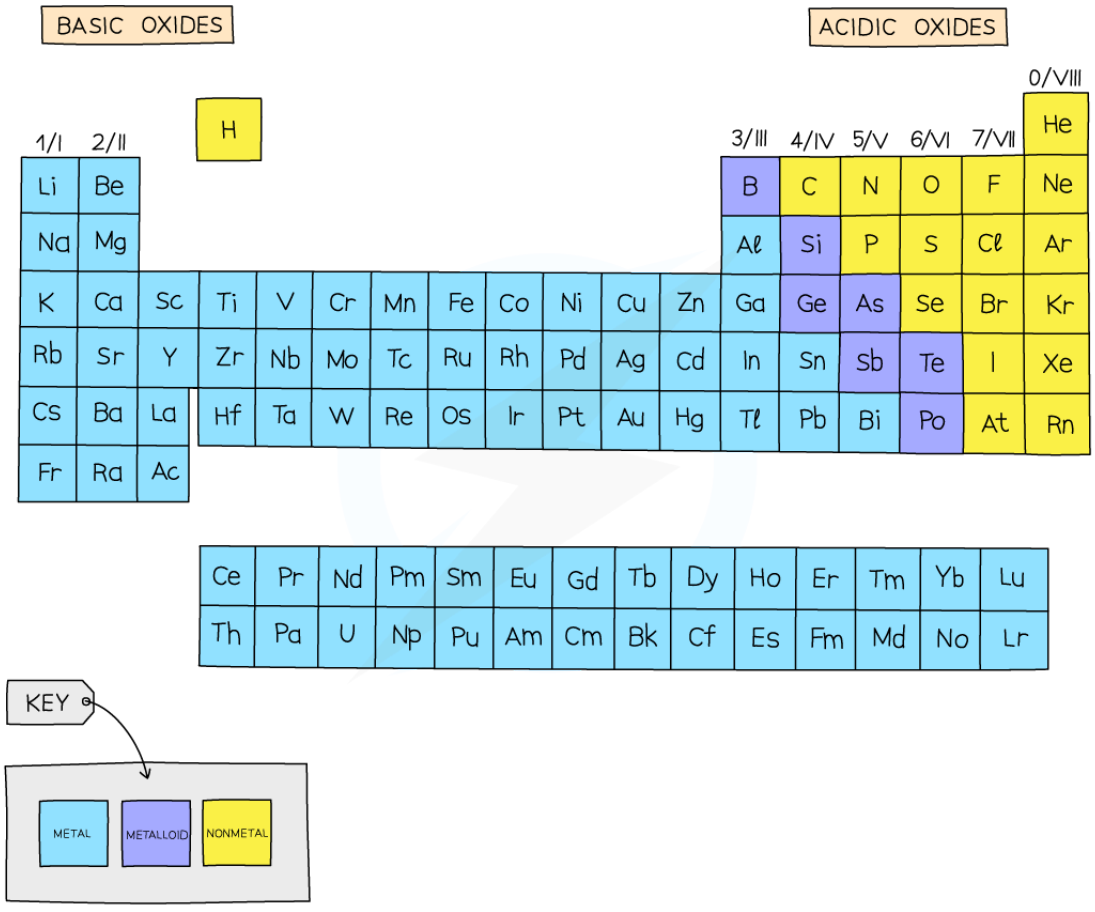
Properties of Metals

Your notes

Physical Properties of Metals & Non-Metals

Metals and non-metals

- The Periodic Table contains over 100 different elements
- They can be divided into two broad types: **metals** and **non-metals**
- Most of the elements are metals and a small number of elements display properties of both types
 - These elements are called **metalloids** or **semimetals**



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The metallic character diminishes moving left to right across the Periodic Table

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Properties of metals

- Conduct **heat** and **electricity**
- Are **malleable** (can be hammered and made into different shapes) and **ductile** (can be drawn into wires)
- Tend to be **lustrous** (shiny)
- Have **high** density and usually have **high** melting points
- Form **positive** ions through electron loss
- Form **basic oxides**

Properties of non-metal elements

- Do **not** conduct heat and electricity
- Are **brittle** when solid and easily break up
- Tend to be dull and nonreflective
- Have **low** density and **low** melting points (many are gases at room temperature)
- Form **negative** ions through electron gain (except for hydrogen)
- Form **acidic oxides**


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Chemical Properties of Metals

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General chemical properties of metals

- The chemistry of metals is studied by analysing their reactions with water, dilute acid and oxygen
- Based on these reactions, a reactivity series of metals can be produced

Reactivity with water

- Some metals react with water, either warm or cold, or with steam
- Metals that react with cold water form a metal hydroxide and hydrogen gas

metal + water → metal hydroxide + hydrogen

- For example calcium:

$$Ca\,(s) + 2H_2O\,(l) \rightarrow Ca(OH)_2\,(aq) + H_2\,(g)$$

- Metals that react with steam form metal oxide and hydrogen gas, for example zinc:

$$Zn\,(s) + H_2O\,(g) \rightarrow ZnO\,(s) + H_2\,(g)$$

Reactivity with acids

- Most metals react with dilute acids such as HCl
- When acids and metals react, the hydrogen atom in the acid is replaced by the metal atom to produce a salt and hydrogen gas, for example iron:

metal + acid → salt + hydrogen

$$Fe\,(s) + 2HCl\,(aq) \rightarrow FeCl_2\,(aq) + H_2\,(g)$$

Reactivity with oxygen

- Unreactive metals such as gold and platinum do not react with oxygen
- Some reactive metals such as the alkali metals react easily with oxygen
- Copper and iron can also react with oxygen although much more slowly
- When metals react with oxygen a metal oxide is formed, for example copper:

metal + oxygen → metal oxide

$$2Cu\,(s) + O_2\,(g) \rightarrow 2CuO\,(s)$$

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9.2 Uses of Metals

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Uses of Metals

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Uses of Metals

Your notes

Uses of Metals

Uses of Aluminium

Use	Most important property
Aeroplane bodies	High strength–to–weight ratio, low density
Overhead power cables	Good conductor of electricity, low density
Saucepans	Good conductor of heat
Food cans	Non–toxic, resistant to corrosion and acidic food stuffs
Window frames	Resistant to corrosion

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Uses of Copper

USE	MOST IMPORTANT PROPERTY
ELECTRICAL WIRES	GOOD CONDUCTOR OF ELECTRICITY AND MALLEABLE
WATER PIPES	EASY TO WORK WITH AND BEND, NON–TOXIC AND UNREACTIVE (DOES NOT REACT WITH WATER)

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9.3 Alloys & their Properties

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* Alloys

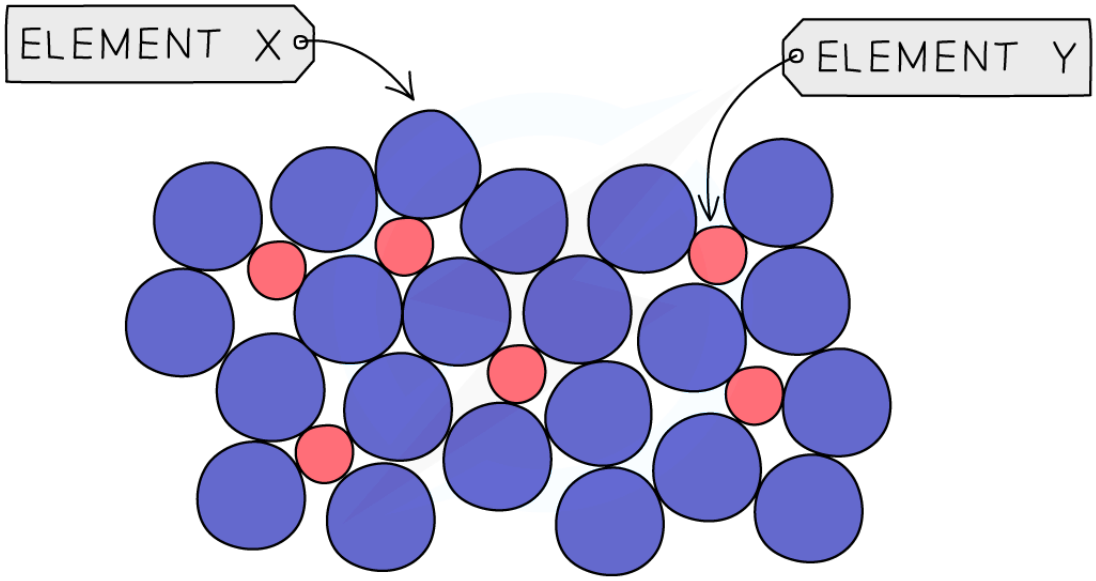
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Alloys

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Properties & Uses of Alloys

- An alloy is a **mixture** of two or more metals or metal with a non-metal such as carbon
- Alloys often have **properties** that can be very **different** from the metals they contain, for example, they can have more **strength**, **hardness** or **resistance** to **corrosion** or extreme **temperatures**
- These enhanced properties can make alloys more useful than pure metals



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The regular arrangement of a metal lattice structure is distorted in alloys

Common alloys and their uses

- **Brass** is an alloy of copper and zinc and is much **stronger** than either metal
 - It is used in musical instruments, ornaments and door knobs
- **Stainless steel** is a mixture of iron and other elements, for example, chromium, nickel and carbon
 - It is used in cutlery because of its hardness and resistance to corrosion
- Alloys of iron with tungsten are extremely **hard** and resistant to high **temperatures**
- Alloys of iron mixed with chromium or nickel are resistant to **corrosion**
- Aluminium is mixed with copper, manganese and silicon for aircraft body production as the alloy is **stronger** but still has a low **density**

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Exam Tip

Alloys are mixtures of substances, they are not chemically combined and an alloy is not a compound.

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Explaining the Properties of Alloys

- Alloys contain atoms of **different** sizes, which **distorts** the normally regular arrangements of atoms in metals
- This makes it more difficult for the layers to **slide** over each other, so alloys are usually much harder than the pure metal

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9.4 Reactivity Series

Contents

- Reactivity Series
- Explaining Reactivity

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Reactivity Series


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Reactivity Series

- The chemistry of the metals is studied by analysing their reactions with water and acids
- Based on these reactions a reactivity series of metals can be produced
- The series can be used to place a group of metals in **order of reactivity** based on the observations of their reactions with water and acids
- The non-metals hydrogen and carbon are also included in the reactivity series as they are used to extract metals from their oxides

Table of Metal Reactions

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Metal	Reaction with cold water	Reaction with acid	Reaction with oxygen
Most reactive			
Potassium	Reacts violently	Reacts violently	Reacts quickly in air
Sodium	Reacts quickly	Reacts violently	Reacts quickly in air
Calcium	Reacts less strongly	Reacts vigorously	Reacts readily
Magnesium	Slow reaction (reacts with steam)	Reacts vigorously	Reacts readily
Aluminium	Slow reaction (reacts with steam)	Reacts readily	Reacts readily
Carbon			
Zinc	Very slow reaction (reacts slowly with steam)	Reacts less strongly	Reacts
Iron	Very slow reaction (reacts slowly with steam)	Reacts less strongly	Reacts
Hydrogen			
Copper	No reaction with water or steam	No reaction	Reacts
Silver	No reaction with water or steam	No reaction	Reacts
Gold	No reaction with water or steam	No reaction	No reaction
Least reactive			

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The reactivity series mnemonic

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- Observations from the table above allow the following reactivity series to be deduced
- The order of this reactivity series can be memorised using the following mnemonic
 - "Please send cats, monkeys and cute zebras into hot countries signed Gordon"

Your notes

Metal	Abbreviation
Most reactive	
Potassium	P – Please
Sodium	S – Send
Calcium	C – Cats,
Magnesium	M – Monkeys,
Aluminium	A – And
Carbon	C – Cute
Zinc	Z – Zebras
Iron	I – Into
Hydrogen	H – Hot
Copper	C – Countries
Silver	S – Signed
Gold	G – Gordon
Least reactive	

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You can learn the reactivity series with the help of a silly phrase

Reactions of Metals

Reaction with cold water

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- The more reactive metals will react with cold water to form a metal hydroxide and hydrogen gas
- Potassium, sodium and calcium all undergo reactions with cold water as they are the most reactive metals:

$$\text{metal} + \text{water} \rightarrow \text{metal hydroxide} + \text{hydrogen}$$

- For example, calcium and potassium:

$$Ca\,(s) + 2H_2O\,(l) \rightarrow Ca(OH)_2\,(aq) + H_2\,(g)$$

$$K\,(s) + H_2O\,(l) \rightarrow KOH\,(aq) + H_2\,(g)$$

Reaction with steam

- Metals just below calcium in the reactivity series do not react with cold water but will react with steam to form a metal oxide and hydrogen gas, for example, magnesium:

$$Mg\,(s) + H_2O\,(g) \rightarrow MgO\,(s) + H_2\,(g)$$

Reaction with dilute acids

- Only metals **above hydrogen** in the reactivity series will react with dilute acids
- Unreactive metals below hydrogen, such as gold, silver and copper, do not react with acids
- The more reactive the metal then the more vigorous the reaction will be
- Metals that are placed high on the reactivity series such as potassium and sodium are very dangerous and react **explosively** with acids
- When acids react with metals they form a **salt** and **hydrogen gas:**
- The general equation is:

metal + acid ⟶ salt + hydrogen

- Some examples of metal-acid reactions and their equations are given below:

Acid-Metal Reactions Table

Acid	Sulfuric Acid	Hydrochloric Acid
Magnesium	$Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$	$Mg + 2HCl \longrightarrow MgCl_2 + H_2$
Zinc	$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$	$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$
Iron	$Fe + H_2SO_4 \longrightarrow FeSO_4 + H_2$	$Fe + 2HCl \longrightarrow FeCl_2 + H_2$

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Reaction with oxygen

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- Some reactive metals, such as the alkali metals, react easily with oxygen
- Silver, copper and iron can also react with oxygen although much more slowly
- When metals react with oxygen a metal oxide is formed, for example, copper:

metal + oxygen → metal oxide

$$2Cu\,(s) + O_2\,(g) \rightarrow 2CuO\,(s)$$

- Gold does not react with oxygen

Deducing the order of reactivity

- The order of reactivity of metals can be deduced by making experimental observations of reactions between metals and water, acids and oxygen
- The more vigorous the reaction of the metal, the higher up the reactivity series the metal is
- A combination of reactions may be needed, for example, the order of reactivity of the more reactive metals can be determined by their reactions with water
- The less reactive metals react slowly or not at all with water, so the order of reactivity would need to be determined by observing their reactions with dilute acid
- Temperature change in a reaction can also be used to determine the order of reactivity
- The greater the temperature change in a reaction involving a metal, the more reactive the metal is


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Explaining Reactivity


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Explaining Reactivity

- Metal atoms form **positive ions** by loss of electrons when they react with other substances
- The **tendency** of a metal to lose electrons is a measure of how reactive the metal is
- A metal that is high up on the series loses electrons easily and is thus **more reactive** than one which is lower down on the series

Displacement reactions between metals and aqueous solutions of metal salts

- Any metal will displace another metal that is **below** it in the reactivity series from a solution of one of its salts
- This is because more reactive metals lose electrons and form ions more readily than less reactive metals, making them better **reducing agents**
- The less reactive metal is a better electron acceptor than the more reactive metal, thus the less reactive metal is reduced. (OIL-RIG: reduction is gain of electrons)l

Example: Magnesium + copper sulfate

- Magnesium is a reactive metal and can displace copper from a copper sulfate solution
- Magnesium loses its electrons more easily and the ion of the less reactive metal, copper, will gain these electrons to form elemental copper
- This is easily seen as the more reactive metal slowly **disappears** from the solution, **displacing** the less reactive metal

$$\text{magnesium} + \text{copper sulfate} \rightarrow \text{magnesium sulfate} + \text{copper}$$

$$Mg\,(s) + CuSO_4\,(aq) \rightarrow MgSO_4\,(aq) + Cu\,(s)$$

- The blue colour of the $CuSO_4$ solution **fades** as colourless magnesium sulfate solution is formed
- Copper coats the surface of the magnesium and also forms solid metal which falls to the bottom of the beaker

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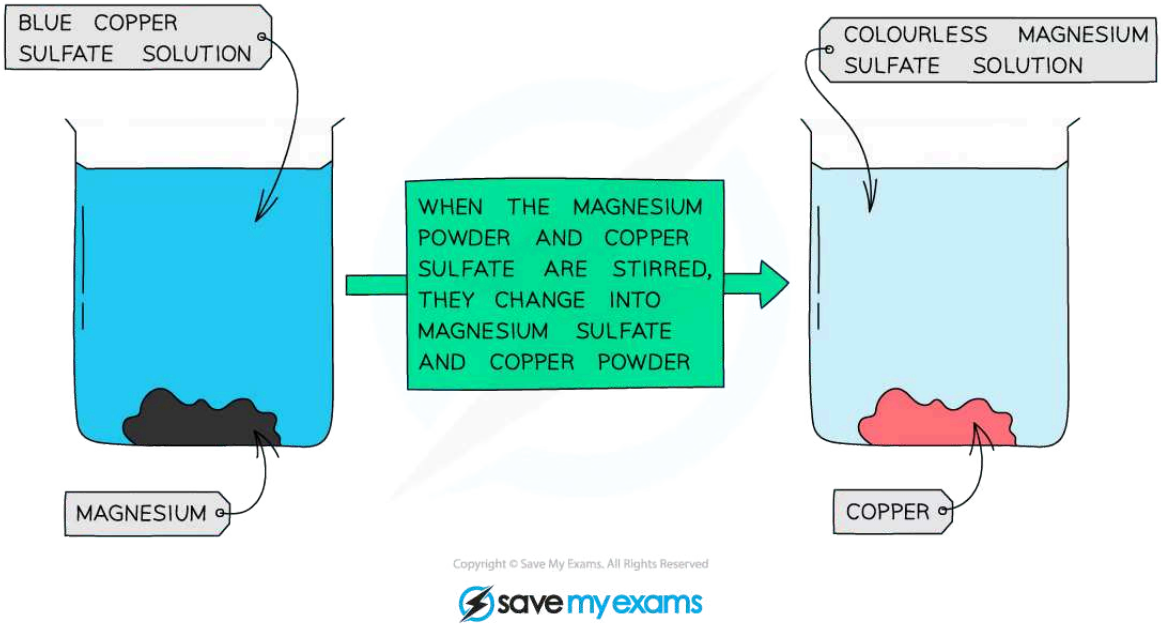


Diagram showing the colour change when magnesium displaces copper from copper sulfate

- By combining different metals and metal salts solutions it is possible to come up with a relative reactivity order

Metal Solutions Displacement Table

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Your notes

Mixture	Products	Equation for reaction
Magnesium and iron(II) sulfate	Magnesium sulfate and iron	$Mg + FeSO_4 \longrightarrow MgSO_4 + Fe$
Zinc and iron chloride	Zinc chloride and iron	$Zn + FeCl_2 \longrightarrow ZnCl_2 + Fe$
Iron and silver nitrate	Iron(II) nitrate and silver	$Fe + 2AgNO_3 \longrightarrow Fe(NO_3)_2 + 2Ag$
Copper and iron(II) chloride	No reaction as iron is more reactive than copper	–
Silver and copper(II) sulfate	No reaction as copper is more reactive than silver	–
Zinc and magnesium chloride	No reaction as magnesium is more reactive than zinc	–

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- From this table we can deduce the order of reactivity:
 - Magnesium and zinc are both more reactive than iron but magnesium is more reactive than zinc
 - Copper and silver are both less reactive than iron but silver is less reactive than copper
- The order of reactivity of the metals tested can be therefore be deduced as:
 - Mg > Zn > Fe > Cu > Ag

Reactivity of aluminium

- Aluminium is high in the reactivity series, but in reality, it does not react with water and the reaction with dilute acids can be quite slow
- This is because it reacts readily with oxygen, forming a protective layer of aluminium oxide which is very thin
- This layer prevents reaction with water and dilute acids, so aluminium can behave as if it is unreactive

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9.5 Corrosion of Metals

Contents

- Rusting of Iron
- Galvanising & Sacrificial Protection

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Rusting of Iron

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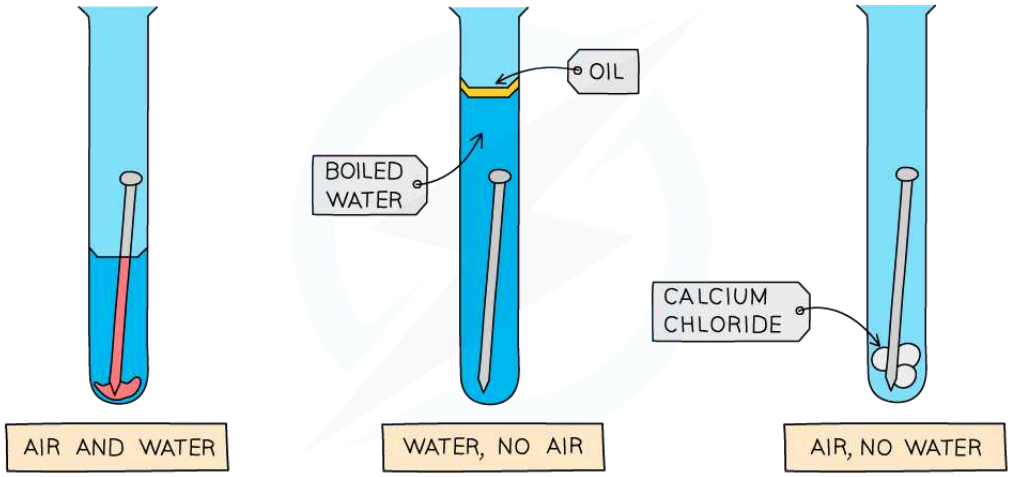
Rusting of Iron

- Rust is a chemical reaction between iron, water and oxygen to form the compound hydrated iron(III) oxide (rust)
- Oxygen and water must be present for rust to occur
- During rusting, iron is oxidised

$$\text{iron} + \text{water} + \text{oxygen} \rightarrow \text{hydrated iron(III) oxide}$$

Investigating rusting

- To investigate the conditions required for rusting, prepare three test tubes as shown in the diagram
- The oil in the 2nd tube keeps out air and the water has been boiled so that no air is left in it
- The calcium chloride in the 3rd tube is used to remove any moisture in the air
- After a few days, the iron nail in the 1st tube will be the only nail to show signs of rust



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Diagram showing the requirements of oxygen and water for rust to occur: only the nail on the left rusts

Rust prevention methods

Barrier methods

- Rust can be prevented by coating iron with barriers that prevent the iron from coming into contact with water and oxygen

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- However, if the coatings are washed away or scratched, the iron is once again exposed to water and oxygen and will rust

Your notes

COMMON BARRIER METHODS	
GREASE	OIL
PAINT	PLASTIC

Exam Tip

Only iron or steel (an alloy made from iron) can **rust**. If any other metal oxidises in air causing the metal to break down, you should say that the metal has **corroded**.

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Galvanising & Sacrificial Protection

Your notes

Galvanising & Sacrificial Protection

- Iron can be prevented from rusting using the **reactivity series**

Sacrificial Protection

- A **more** reactive metal can be attached to a **less** reactive metal
- The more reactive metal will oxidise and therefore corrode first, protecting the less reactive metal from corrosion
- E.g. using zinc bars on the side of steel ships:

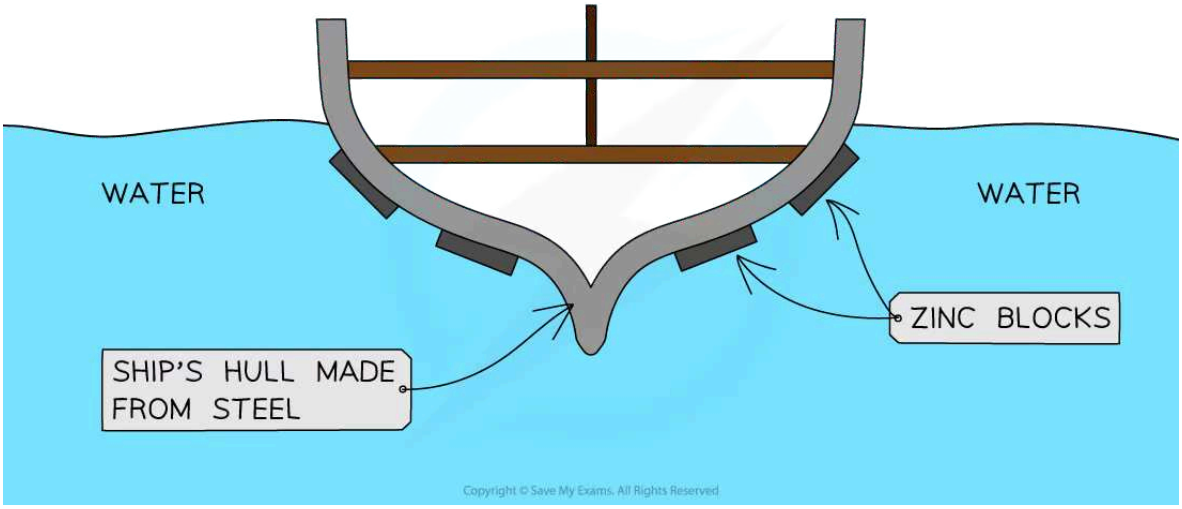


Diagram to show the use of zinc bars on the sides of steel ships as a method of sacrificial protection

- Zinc is more reactive than iron therefore will lose its electrons more easily than iron and is oxidised more easily:

$$Zn \rightarrow Zn^{2+} + 2e^-$$

- The iron is less reactive therefore will not lose its electrons as easily so it is not oxidised; the zinc is **sacrificed** to protect the steel
- For continued protection, the zinc bars have to be replaced before they completely corrode

Galvanising

- **Galvanising** is a process where the iron to be protected is coated with a layer of zinc
- This can be done by electroplating or dipping it into molten zinc
- $ZnCO_3$ is formed when zinc reacts with oxygen and carbon dioxide in the air and protects the iron by the barrier method

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- If the coating is damaged or scratched, the iron is still protected from rusting by **sacrificial protection**


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Exam Tip

You may be asked to explain why a metal is/is not suitable as a method of preventing an iron/steel object from rusting.

Remember that if it is higher in the reactivity series than iron, it will be suitable for sacrificial protection as it will be oxidised instead of iron.

If it is lower in the reactivity series than iron, it would not be suitable as iron would be oxidised, causing it to rust.

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9.6 Extraction of Metals

Contents

- Extraction of Metals
- Extraction of Iron from Hematite
- Extraction of Aluminium from Bauxite

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Extraction of Metals

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Extraction of Metals

- The Earth's crust contains metals and metal compounds such as gold, copper, iron oxide and aluminium oxide
- Useful metals are often **chemically combined** with other substances forming **ores**
- A metal ore is a rock that contains enough of the metal to make it worthwhile extracting
- They have to be extracted from their ores through processes such as **electrolysis**, using a **blast furnace** or by reacting with **more reactive** material
- In many cases the ore is an **oxide** of the metal, therefore the extraction of these metals is a reduction process since oxygen is being removed
- Common examples of oxide ores are **iron** and **aluminium** ores which are called **hematite** and **bauxite** respectively
- Unreactive metals do not have to be extracted chemically as they are often found as the **uncombined** element
- This occurs as they do not easily react with other substances due to their **chemical stability**
- They are known as **native metals** and examples include **gold** and **platinum** which can both be mined directly from the Earth's crust
- The position of the metal on the reactivity series influences the method of extraction
- Those metals placed higher up on the series (above carbon) have to be extracted using electrolysis
- Metals lower down on the series can be extracted by heating with carbon

The Extraction Method Depends on the Position of a Metal in the Reactivity Series

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Metal	Method of extraction
Most reactive	
Potassium	Extracted by electrolysis of the molten chloride or molten oxide Large amount of electricity required so expensie process
Sodium	
Lithium	
Calcium	
Magnesium	
Aluminium	
Carbon	
Zinc	Extracted by heating with a reducing agent such as carbon or carbon monoxide in a blast furnace Cheap process as carbon is cheap and can be source of heat as well Found as pure elements
Iron	
Hydrogen	
Copper	
Silver	
Gold	
Least reactive	

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Extraction of Iron from Hematite

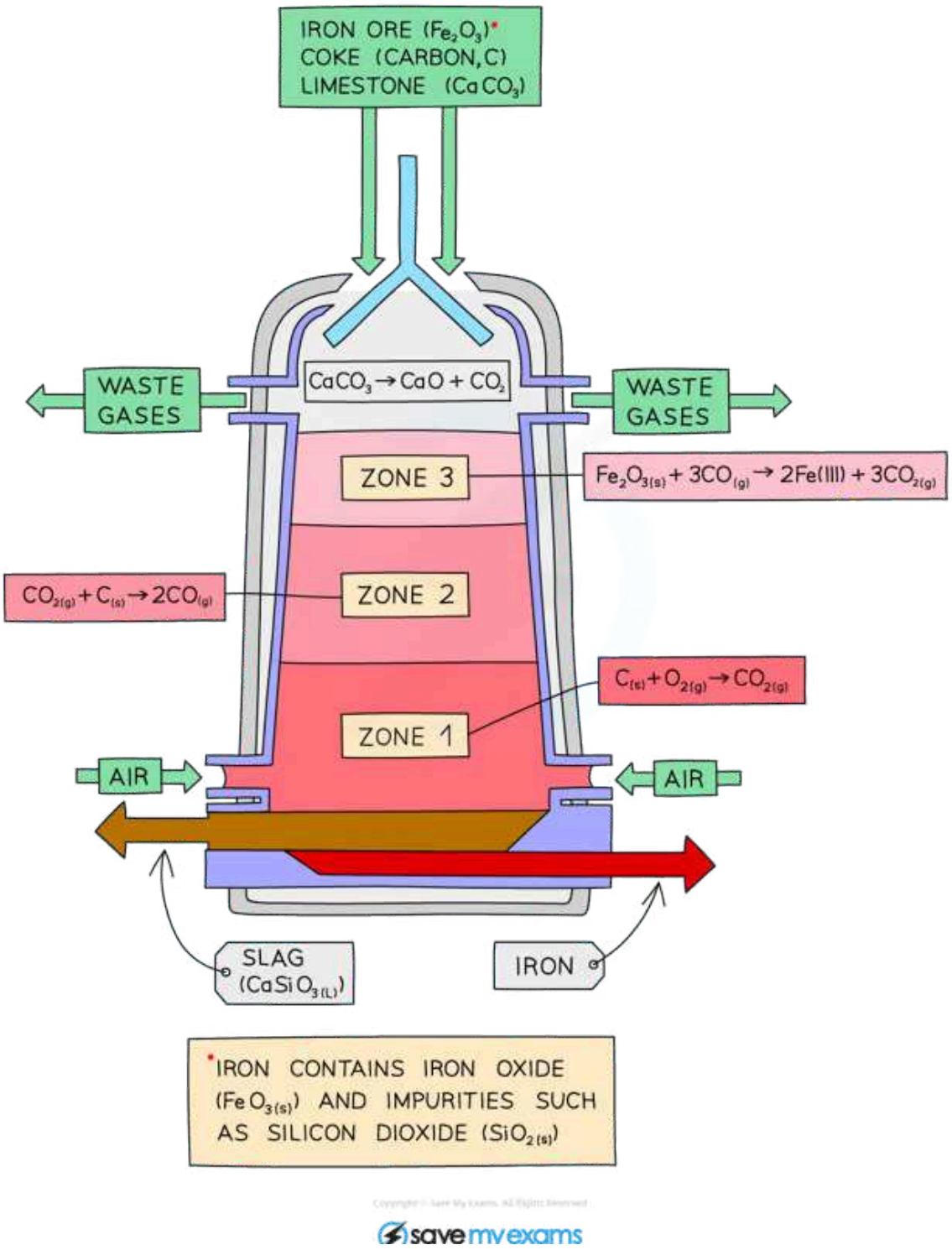

Your notes

Extraction of Iron from Hematite

- Iron is extracted in a large container called a blast furnace from its ore, hematite
- Modern blast furnaces produce approximately 10,000 tonnes of iron per day
- The process is demonstrated and explained below:

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Your notes

Diagram showing the carbon extraction of iron

- The raw materials: iron ore (hematite), coke (an impure form of carbon), and limestone are added into the top of the blast furnace
- Hot air is blown into the bottom
- **Zone 1:**
 - Coke **burns** in the hot air forming **carbon dioxide**
 - The reaction is exothermic so it gives off heat, **heating** the furnace

carbon + oxygen → carbon dioxide

- **Zone 2**:
 - At the high temperatures in the furnace, more coke reacts with carbon dioxide forming carbon monoxide
 - Carbon dioxide has been **reduced** to carbon monoxide

carbon + carbon dioxide → carbon monoxide

- **Zone 3**:
 - Carbon monoxide **reduces** the iron(III) oxide in the iron ore to form iron
 - This will melt and collect at the bottom of the furnace, where it is tapped off:

iron(III) oxide + carbon monoxide → iron + carbon dioxide

- Limestone (calcium carbonate) is added to the furnace to remove impurities in the ore.
 - The calcium carbonate in the limestone **thermally decomposes** to form calcium oxide

calcium carbonate → calcium oxide + carbon dioxide

- The calcium oxide formed reacts with the silicon dioxide, which is an impurity in the iron ore, to form calcium silicate
- This melts and collects as a molten **slag** floating on top of the molten iron, which is tapped off separately

calcium oxide + silicon dioxide → calcium silicate

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Equations for Extraction of Iron from Hematite

Your notes

- The symbol equations for the different stages of the extraction of iron from hematite are:
- **Zone 1**: The burning of carbon (coke) to provide heat and produce carbon dioxide:

$$C\,(s) + O_2\,(g) \rightarrow CO_2\,(g)$$

- **Zone 2**: The reduction of carbon dioxide to carbon monoxide:

$$CO_2\,(g) + C\,(s) \rightarrow 2CO\,(g)$$

- **Zone 3**: The reduction of iron(III) oxide by carbon monoxide:

$$Fe_2O_3\,(s) + 3CO\,(g) \rightarrow 2Fe\,(l) + 3CO_2\,(g)$$

- The thermal decomposition of calcium carbonate (limestone) to produce calcium oxide:

$$CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g)$$

- The formation of slag:

$$CaO\,(s) + SiO_2\,(s) \rightarrow CaSiO_3\,(l)$$

Exam Tip

You must learn each of these equations!

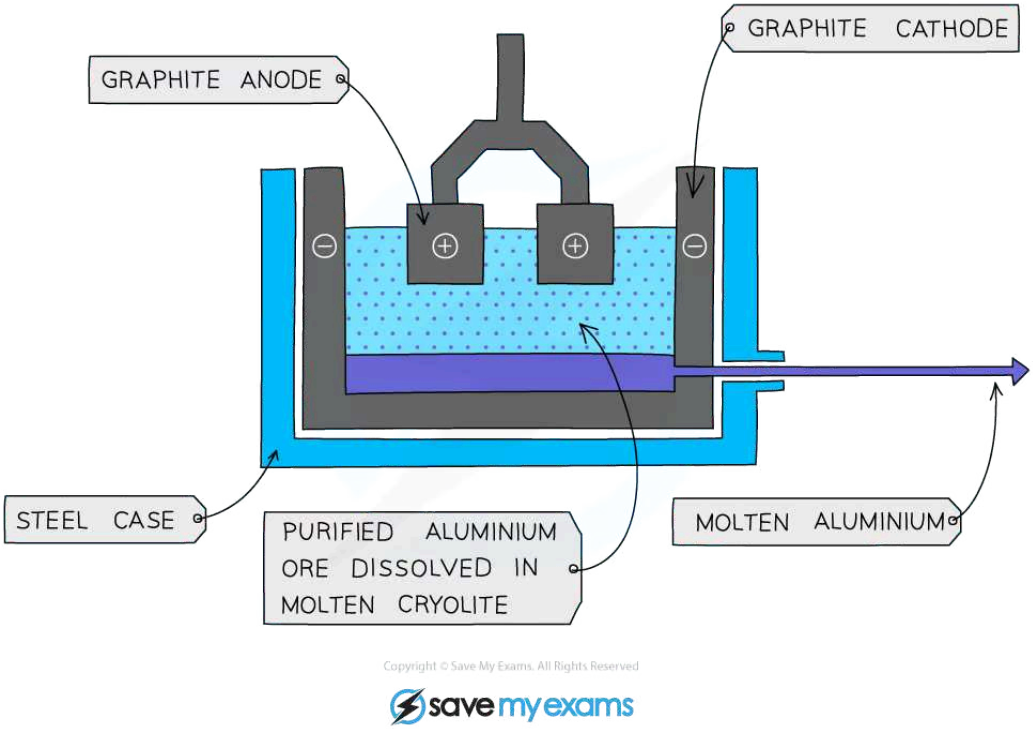
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Extraction of Aluminium from Bauxite

Your notes

Extraction of Aluminium from Bauxite

- Aluminium is a reactive metal, above carbon in the reactivity series
- Its main ore, is **bauxite**, which contains aluminium oxide
- Aluminium is higher in the reactivity series than carbon, so it cannot be extracted by reduction using carbon
- Instead, aluminium is extracted by **electrolysis**



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Diagram showing the extraction of aluminium by electrolysis

- Bauxite is first **purified** to produce aluminium oxide, Al_2O_3
- Aluminium oxide is then dissolved in **molten cryolite**
 - This is because aluminium oxide has a melting point of over 2000°C which would use a lot of energy and be very expensive
 - The resulting mixture has a lower melting point without interfering with the reaction
- The mixture is placed in an electrolysis cell, made from steel, lined with graphite
- The **graphite** lining acts as the negative electrode, with several large graphite blocks as the positive electrodes

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- At the **cathode** (negative electrode):
 - Aluminium ions gain electrons (reduction)
 - Molten aluminium forms at the bottom of the cell
 - The molten aluminium is siphoned off from time to time and fresh aluminium oxide is added to the cell

$$Al^{3+} + 3e^- \rightarrow Al$$

- At the **anode** (positive electrode):
 - Oxide ions lose electrons (oxidation)
 - Oxygen is produced at the anode:

$$2O^{2-} \rightarrow O_2 + 4e^-$$

- The overall equation for the reaction is:

$$2Al_2O_3 \rightarrow 4Al + 3O_2$$

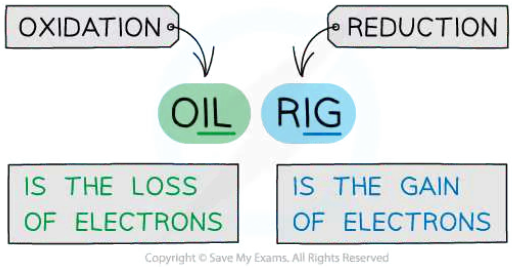
- The carbon in the **graphite** anodes reacts with the oxygen produced to produce **CO_2**

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

- As a result the anode wears away and has to be replaced regularly
- A lot of **electricity** is required for this process of extraction, this is a major **expense**

Exam Tip

Use OIL RIG to remember whether oxidation or reduction has occurred at the electrodes:



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Your notes

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Cambridge O Level Chemistry

Your notes

10.1 Water & Water Pollution

Contents

- Water: Chemical Tests
- Substances in Water from Natural Sources
- Water Treatment

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Water: Chemical Tests

Your notes

Chemical Tests for Water

Chemical tests for water

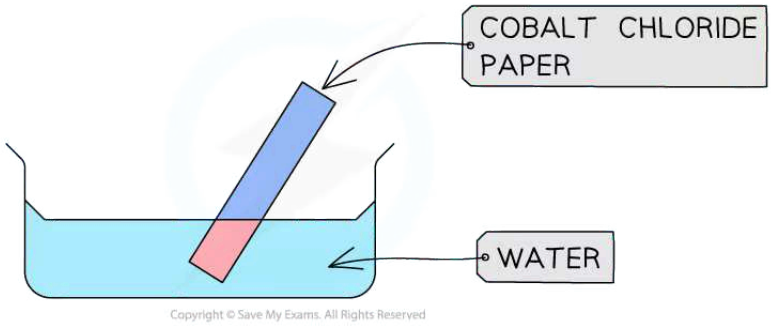
Using cobalt(II) chloride:

- Cobalt(II) chloride turns **blue** to **pink** on the addition of water. This test is usually done using cobalt chloride paper
- The equation is:

anhydrous cobalt(II) chloride + water ⇌ hydrated cobalt(II) chloride

$$CoCl_2(s) + 6H_2O(l) \rightleftharpoons CoCl_2.6H_2O(s)$$

blue pink



Test for water using cobalt chloride paper which turns pink in the presence of water

Using copper(II) sulfate:

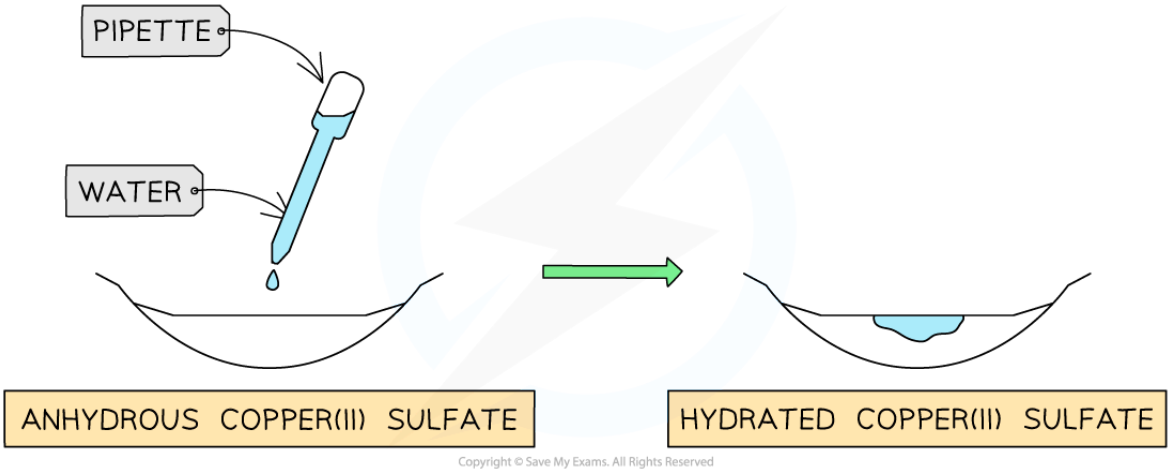
- Anhydrous copper(II) sulfate turns **white** to **blue** on the addition of water
- The equation is:

anhydrous copper(II) sulfate + water ⇌ hydrated copper(II) sulfate

$$CuSO_4(s) + 5H_2O(l) \rightleftharpoons CuSO_4.5H_2O(s)$$

white blue

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Test for water using anhydrous copper(II) sulfate which turns blue in the presence of water

Exam Tip

Make sure you learn the symbol equations.

Purity of Water

Testing for purity

- Pure substances melt and boil at specific and sharp temperatures
- Water has a boiling point of 100 °C and a melting point of 0 °C
- Mixtures have a range of melting and boiling points as they consist of different substances that melt or boil at different temperatures
- Melting and boiling points data can therefore be used to determine the purity of water
- Impurities tend to increase the boiling point of water, so impure water will boil at temperatures above 100 °C
- Impurities tend to decrease the melting point of water, so impure water will melt at temperatures below 0 °C

Distilled water

- Distilled water is water that has been heated to form a vapour, and then condensed back to a liquid
- It contains very few impurities
- Distilled water is used in practical chemistry because of its high purity
- Tap water contains more impurities which could interfere with chemical reactions so is typically not used

Your notes

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Substances in Water from Natural Sources


Your notes

Substances in Water from Natural Sources

- We use water in many aspects of our everyday life:
 - Domestic uses: for drinking, cooking, gardening and general sanitation
 - Agricultural uses: as a drink for animals and watering crops
 - Industrial uses: as a solvent, as a coolant and heated to make steam used to generate electricity
- Water is found in natural sources such as lakes, rivers and underground water sources (groundwater)
- A rock that stores water is known as an **aquifer**
- Water from natural sources may contain a variety of different substances, including:
 - Dissolved oxygen
 - Metal compounds
 - Plastics
 - Sewage
 - Harmful microbes
 - Nitrates from fertilisers
 - Phosphates from fertilisers and detergents
- Many of these substances enter water sources when rain falls and washes them into lakes, rivers or groundwater
- Some of these substances are naturally occurring but many are a direct result of human activities

Exam Tip

Despite the term 'natural source', water from these sources may contain lots of impurities and harmful substances and should not be confused with water which is clean and ready for use.

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Beneficial & Harmful Substances in Water


Your notes

- Some of the substances which may be found in natural water sources are beneficial and others may have harmful effects
- Beneficial substances include:
 - Dissolved oxygen - essential for aquatic life
 - Metal compounds - some provide essential minerals which are necessary for life, such as calcium and magnesium
- Potentially harmful substances include:
 - Metal compounds - some are toxic like aluminium and lead
 - Some plastics - these may be harmful to aquatic life in many ways, e.g. getting trapped in plastic waste, dying of starvation as their stomach is filled with plastic
 - Sewage - contains harmful microbes which can cause disease
 - Nitrate & phosphates from fertilisers - these can promote the growth of aquatic plant life which leads to deoxygenation of water. Ultimately, this can cause damage to aquatic life in a process called eutrophication (you do not need to know the details of this process)

Exam Tip

Metal compounds can be both beneficial and harmful, it depends on the metal within the compound, e.g. calcium, magnesium, potassium and sodium are all metals which are essential for life. Harmful metals include lead, arsenic and mercury.

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Water Treatment


Your notes

Water Treatment

Water supply

- Water is taken from **rivers**, **reservoirs** or underground water sources (**groundwater**)
- A rock that stores water is known as an **aquifer**

Water treatment

- Untreated water contains **soluble** and **insoluble** impurities
- Insoluble impurities include soil, pieces of plants and other organic matter
- Soluble impurities include dissolved calcium, metallic compounds and inorganic pollutants
- Water is pumped into sedimentation tanks where the water is allowed to stand for a few hours
- Mud, sand and other particles will fall to the bottom of the tank due to gravity and form a layer of sediment, in a process called **sedimentation**
- **Filtration** is the process used to remove smaller particles by passing the water through layers of sand and gravel filters that trap solid particles
- Water can also be passed through **carbon** (in the form of charcoal) to remove tastes and odours
- Bacteria and other microorganisms are too small to be trapped by the filters so **chlorination** is used
- This involves the careful addition of **chlorine** to the water supply which kills bacteria and other unwanted microorganisms
- Cholera and typhoid are examples of bacterial diseases which can arise from the consumption of untreated water

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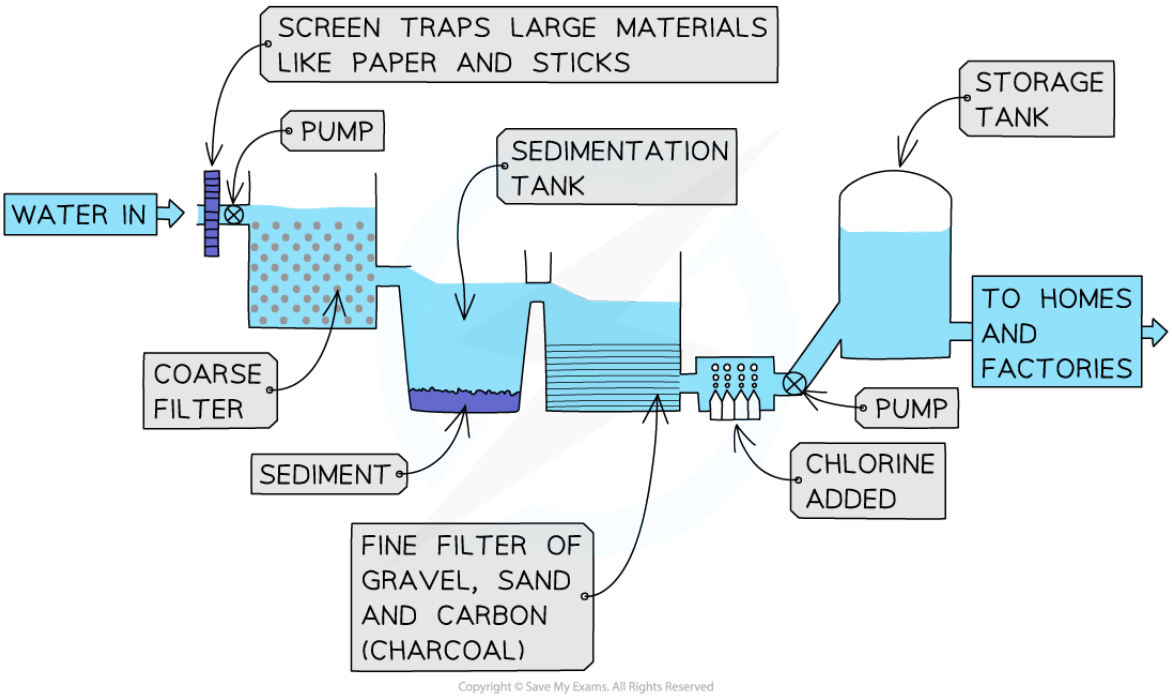


Diagram showing the stages in the treatment of water

Exam Tip

Exam questions on water treatment often focus on the purpose of each stage of the process.

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Cambridge O Level Chemistry

Your notes

10.2 Fertilisers

Contents

Fertilisers

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Fertilisers


Your notes

Fertilisers

N, P, K Fertilisers

- Fertilisers contain **nitrogen, potassium** and **phosphorus**
 - Nitrogen makes chlorophyll and protein and promotes healthy **leaves**
 - Potassium promotes **growth** and healthy **fruit** and **flowers**
 - Phosphorus promotes healthy **roots**
- Fertiliser compounds contain the following water-soluble ions:
 - Ammonium ions, NH_4^+ and nitrate ions, NO_3^-, are sources of soluble nitrogen
 - Phosphate ions, PO_4^{3-} are a source of soluble phosphorus
 - Most common potassium compounds dissolve in water to produce potassium ions, K^+
- Common fertiliser compounds include:
 - Ammonium nitrate, NH_4NO_3
 - Ammonium phosphate, $(NH_4)_3PO_4$
 - Potassium sulfate, K_2SO_4
- Ammonium salts and nitrates are commonly used as fertilisers
- Different fertilisers contain different amounts of fertiliser compounds so each contains different proportions of nitrogen, potassium and phosphorous

Exam Tip

You may be asked to select a compound or combination of compounds from a list which would be the most effective as a fertiliser. Look for the compound or combination of compounds that contain the most elements from nitrogen, phosphorous and potassium.

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Cambridge O Level Chemistry

Your notes

10.3 Air Quality & Climate

Contents

- Air
- Effects of Greenhouse Gases
- Reducing the Effects of Environmental Issues
- Photosynthesis

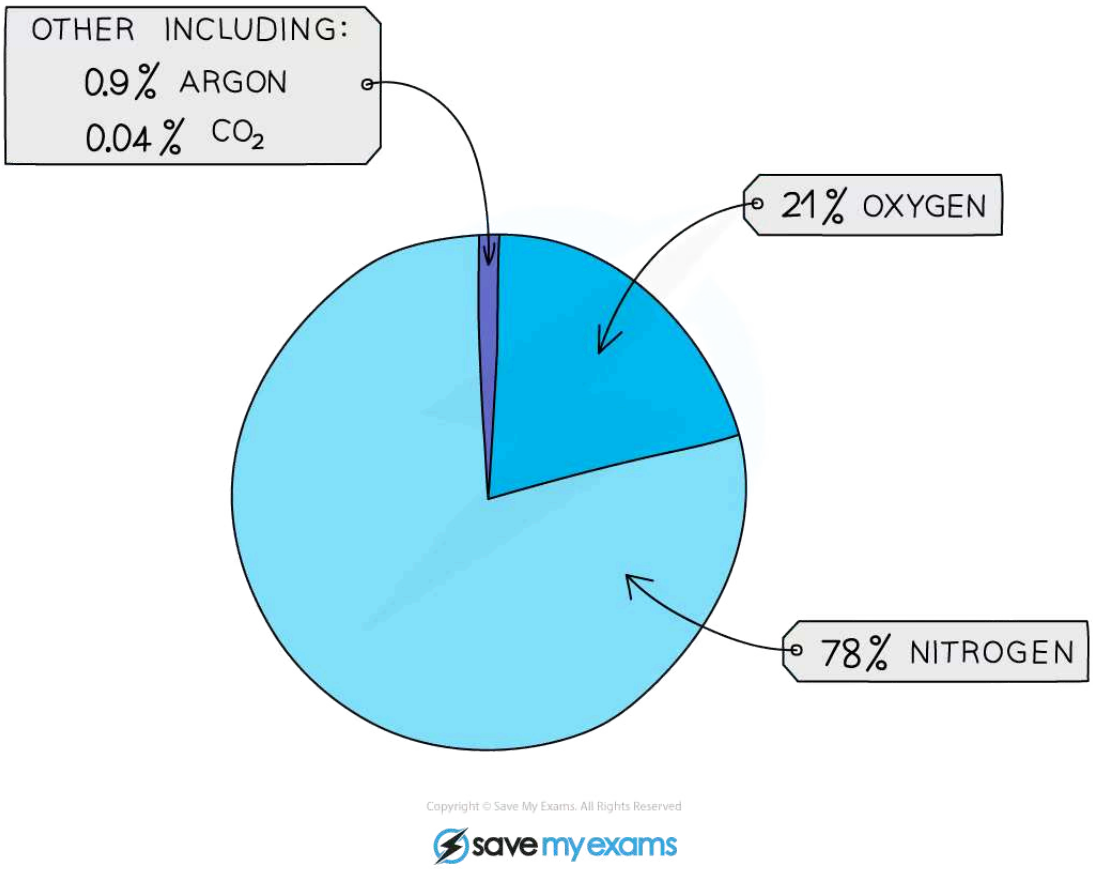
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Air

Your notes

The Composition of Air

- The chart below shows the approximate percentages by volume of the main gases in unpolluted, dry air:



Pie chart showing the composition of clean dry air

Uses of air

- The gases available in the air have many important applications
- The noble gases are used in many applications, e.g. helium is used to fill balloons, argon is used in tungsten light bulbs, krypton is used in lasers for eye surgery
- Oxygen is used in **steel** making, **welding** and **breathing** apparatus

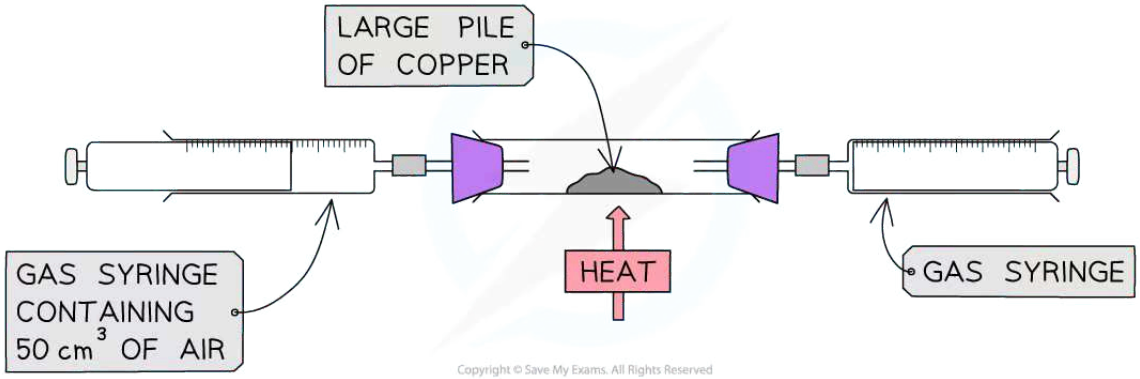
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- Nitrogen is used in **food packaging**, the production of **ammonia** and in the production of **silicon chips**
- Oxygen and nitrogen are separated from the air by fractional distillation

Your notes

Investigating the percentage of oxygen in air

- The percentage of oxygen in the air can be investigated by passing a known quantity of air over a metal
- The oxygen in the air will react with the metal, forming a metal oxide
- The oxygen will be removed from the air and the volume of the air with the oxygen removed can be measured
- An example of the apparatus that can be used to investigate this is shown below:



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Method:

- Heat the copper using a Bunsen burner
- Push the plunger of the syringe containing air, forcing the air into the other plunger until all of the air has transferred
- Push the air back from the now filled plunger to the other plunger
- Repeat this several times for about 3 minutes
- The copper will turn black as copper reacts with the oxygen in the air and copper oxide is produced
- Allow the apparatus to cool
- Ensure all the gas is in one syringe and record the volume of gas
- The percentage of oxygen in the air can be calculated from the results

Worked example

In the experiment above, the starting volume of air was 50.0 cm^3 of air and the final volume of air was 39.5 cm^3

Calculate the percentage of oxygen in the air.

Step 1 - calculate the volume of oxygen in 50.0 cm^3 of air

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volume of oxygen in air = starting volume - final volume

= 50.0 - 39.5 = 10.5 cm^3

Step 2 - calculate the percentage of oxygen in air

$$\text{Percentage of oxygen in air} = \frac{\text{volume of oxygen in air}}{\text{starting volume of air}} \times 100$$

$$= \frac{10.5}{50.0} \times 100$$

= 21.0 %

Your notes

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Air Pollution


Your notes

- In addition to the gases present naturally in our atmosphere, other gases are present due to human activities and are classed as air pollutants

Carbon dioxide

- **Sources:** complete combustion of carbon-containing fuels such as fossil fuels, e.g. the complete combustion of methane:

$$CH_4 + O_2 \rightarrow CO_2 + 2H_2O$$

- **Adverse effects:** increases global warming, which leads to climate change

Carbon monoxide

- **Sources**: incomplete combustion of carbon-containing fuels such as fossil fuels, e.g. incomplete combustion of gasoline:

$$C_8H_{18} + 9O_2 \rightarrow 5CO + 2CO_2 + 9H_2O$$

- **Adverse effects**: toxic, combining with haemoglobin in the blood and prevents it from carrying oxygen

Particulates

- **Sources:** incomplete combustion of carbon-containing fuels such as fossil fuels can also produce particulates of carbon (soot), e.g the incomplete combustion of methane can produce CO and C:

$$2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O$$

$$CH_4 + O_2 \rightarrow C + 2H_2O$$

- **Adverse effects**: respiratory problems and cancer

Methane

- **Sources:** waste gases from digestive processes of animals, decomposition of vegetation, bacterial action in swamps, rice paddy fields and landfill sites
- **Adverse effects**: increases global warming, which leads to climate change

Oxides of nitrogen

- **Sources**: reaction of nitrogen with oxygen in the presence of high temperatures, e.g. in car engines, high-temperature furnaces and when lightning occurs. It is also a product of bacterial action in the soil
- **Adverse effects**:
 - Produces photochemical smog
 - Dissolves in rain to form **acid rain** which causes corrosion to metal structures, buildings and statues made of carbonate rocks, damage to aquatic organisms. Pollutes crops and water supplies, irritates lungs, throats and eyes and causes **respiratory** problems

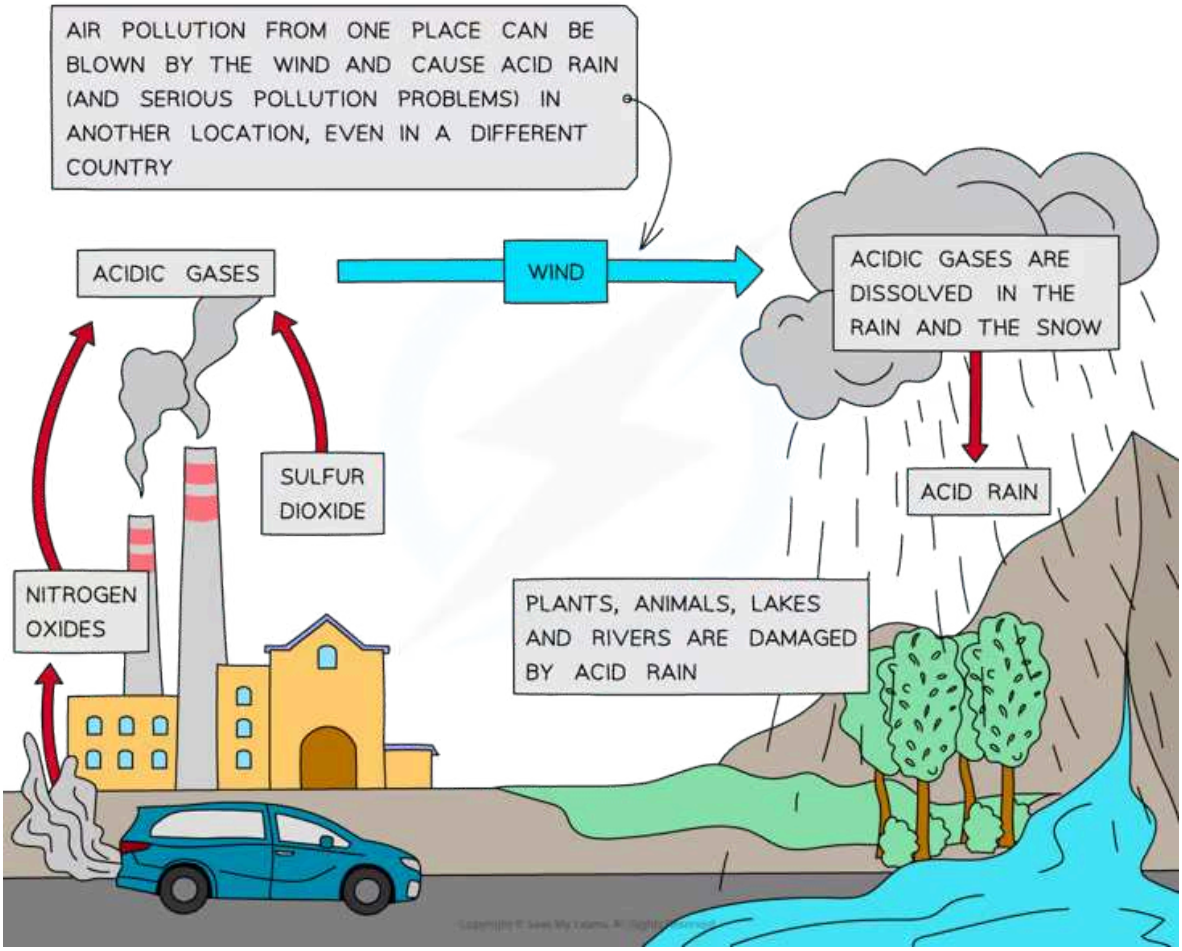
Sulfur dioxide

- **Sources**: combustion of fossil fuels containing sulfur compounds. Power stations are a major source of sulfur dioxide

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- **Adverse effects**: dissolves in rain to form **acid rain** with similar effects as the acid rain caused by oxides of nitrogen

Your notes



How acid rain is produced

Exam Tip

Complete and incomplete combustion of hydrocarbons produce different products. Complete combustion occurs in **excess** oxygen and produces **CO_2** and **H_2O**. Incomplete combustion occurs in **oxygen-deficient** conditions and produces **CO**, **H_2O** and sometimes particulates of **carbon** (soot).

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Effects of Greenhouse Gases

Your notes

Effects of Greenhouse Gases

EXTENDED

- The Sun emits energy in the form of radiation that enters the Earth's atmosphere
- Some thermal energy is reflected from the Earth's surface
- Most thermal energy is absorbed and re-emitted back from the Earth's surface
- The energy passes through the atmosphere where some thermal energy passes straight through and is emitted into space
- But some thermal energy is absorbed by greenhouse gases such as **carbon dioxide** and **methane** and is re-emitted in all directions
- This reduces the thermal energy lost into space and traps it within the Earth's atmosphere, keeping the Earth warm
- This process is known as the **greenhouse effect**
- As the concentration of greenhouse gases in the atmosphere increases due to human activity, more thermal energy is trapped within the Earth's atmosphere causing the Earth's average temperature to rise (global warming)
- This process is called the **enhanced greenhouse effect**

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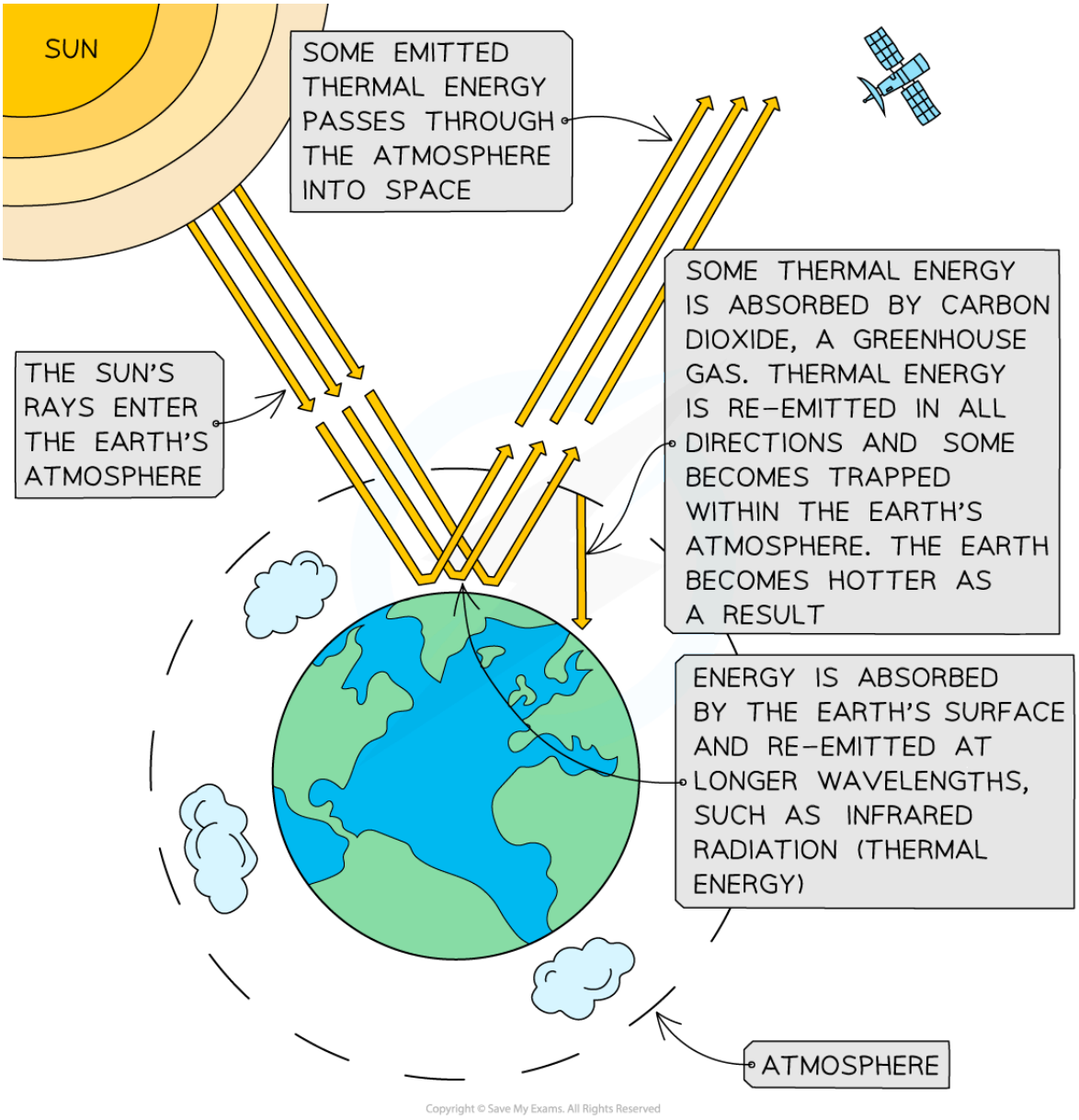


Diagram showing how the greenhouse effect occurs

Consequences of global warming:

- Climate change due to the increase in Earth's temperature

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- Water levels will rise as glaciers melt because of high temperatures, causing flooding in low-lying countries
- Extinction of species due to the destruction of natural habitats
- Migration of species as they will move to areas that are more habitable (no droughts)
- Spread of diseases caused by warmer climate
- Loss of habitat due to climate change (animals that live on glaciers or in low-lying countries)


Your notes

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Reducing the Effects of Environmental Issues


Your notes

Reducing the Effects of Environmental Issues

- The effects of climate change and acid rain are far reaching and strategies are needed to reduce their impact

Dealing with Climate Change

- The production of greenhouse gases needs to be reduced drastically to avoid or at least slow climate change
- CO_2 emissions can be reduced by using hydrogen and renewable energy supplies such as solar or wind energy instead of burning fossil fuels
- Reducing the amount of livestock farming would decrease the methane emissions produced from digestion in animals
- Planting more trees would remove more carbon dioxide from the atmosphere

Dealing with Acid Rain

- Acid rain is caused by oxides of nitrogen and sulfur dioxide
- The effects of acid rain can be reduced by decreasing the amount of oxides of nitrogen and sulfur dioxide that are produced
- Catalytic convertors in vehicles can be used to remove oxides of nitrogen
- Emissions of sulfur dioxide can be reduced by either:
 - Using fuels which contain low levels of sulfur
 - Flue gas desulfurisation - this involves reacting the sulfur dioxide emitted from burning fuels containing sulfur, with calcium oxide therefore removing it from the flue gas

Exam Tip

There are many other ways that carbon dioxide, methane, oxides of nitrogen and sulfur dioxide can be reduced, e.g. by reducing energy usage to reduce CO_2 emissions but it is only the examples stated above that you need to know.

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Oxides of Nitrogen in Car Engines


Your notes

EXTENDED

Oxides of nitrogen

- These compounds (NO and NO_2) are formed when nitrogen and oxygen react in the **high pressure** and **temperature** conditions of internal combustion engines and blast furnaces
- Exhaust gases also contain unburned hydrocarbons and carbon monoxide
- Cars are fitted with catalytic converters which form a part of their exhaust systems
- Their function is to render these exhaust gases harmless

Catalytic converters

- They contain a series of **transition metal catalysts** including platinum and rhodium
- The metal catalysts are in a **honeycomb** within the converter to increase the surface area available for reaction
- A series of redox reactions occurs which neutralises the pollutant gases
- Carbon monoxide is oxidised to carbon dioxide:

$$2CO + O_2 \rightarrow 2CO_2$$

- Oxides of nitrogen are reduced to N_2 gas:

$$2NO \rightarrow N_2 + O_2$$

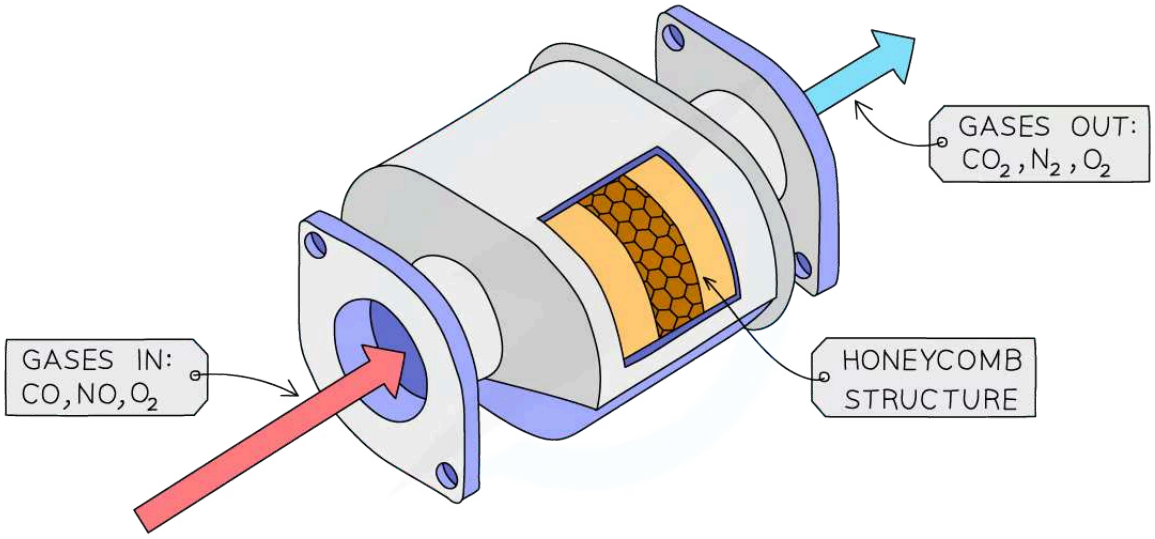
$$2NO_2 \rightarrow N_2 + 2O_2$$

- A single reaction can summarise the reaction of nitrogen monoxide and carbon monoxide within a catalytic convertor:

$$2NO + 2CO \rightarrow N_2 + 2CO_2$$

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Your notes



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Catalytic converters are designed to reduce the polluting gases produced in car exhausts

- Unburned hydrocarbons can also be oxidised to carbon dioxide and water:

$$C_8H_{18} + 12½O_2 \rightarrow 8CO_2 + 9H_2O$$

Exam Tip

Whilst carbon dioxide does not have direct adverse effects to our health, it is a greenhouse gas and is a major contributor to climate change

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Photosynthesis

Your notes

Photosynthesis

Photosynthesis

- Photosynthesis is an endothermic reaction in which energy is transferred from the environment to the chloroplasts in green plants to make glucose
- The reactants for this reaction are carbon dioxide and water
- Glucose and oxygen are produced
- Chlorophyll (found in chloroplasts) and energy from light are required for this reaction to occur
- The word equation for photosynthesis is:



Carbon dioxide and water producing glucose and oxygen during photosynthesis

Symbol Equation for Photosynthesis

EXTENDED

Symbol Equation for Photosynthesis

- The balanced symbol equation for photosynthesis is:



Balanced symbol equation for photosynthesis

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Cambridge O Level Chemistry

Your notes

11.1 Formulae, Functional Groups & Terminology

Contents

- Organic Formulae
- Homologous Series
- Saturated & Unsaturated Compounds

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Organic Formulae

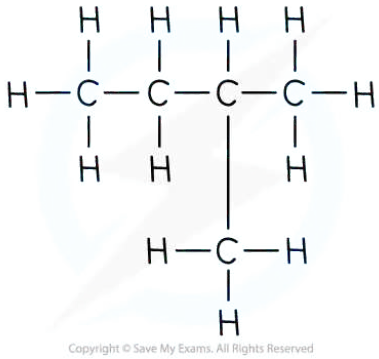
Your notes

Displayed Formulae

- Organic Chemistry is the scientific study of the structure, properties, and reactions of organic compounds.
- Organic compounds are those which contain carbon
- For conventional reasons metal carbonates, carbon dioxide and carbon monoxide are **not** included in organic compounds
- Many of the structures you will be drawing are hydrocarbons
- A **hydrocarbon** is a compound that contains **only** hydrogen and carbon atoms
- Organic compounds can be represented in a number of ways:
 - Displayed Formulae
 - General Formulae
 - Structural Formulae

The **displayed formula** shows the spatial arrangement of all the atoms and bonds in a molecule

- For example:



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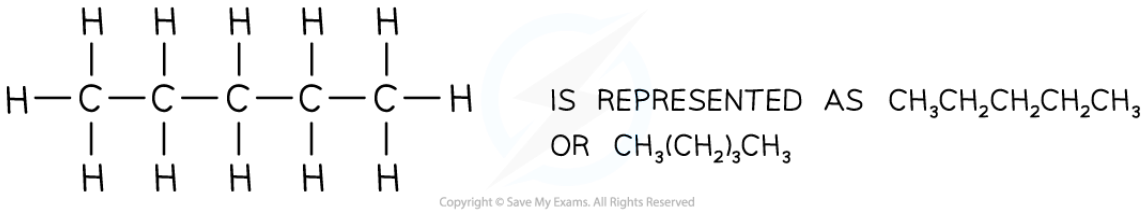
- This displayed formula tells us several things about the compound
 - It has 5 carbon atoms
 - It has 12 hydrogen atoms
 - It has only single bonds

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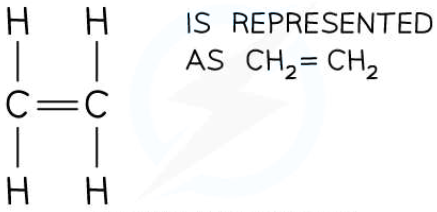
Structural Formulae

Your notes

- In structural formulae, enough information is shown to make the structure clear, but most of the actual covalent bonds are omitted
- Only important bonds are always shown, such as double and triple bonds
- Identical groups can be bracketed together
- Side groups are also shown using brackets
- Straight chain alkanes are shown as follows:



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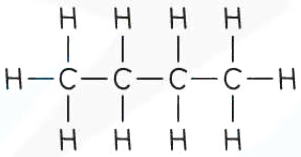
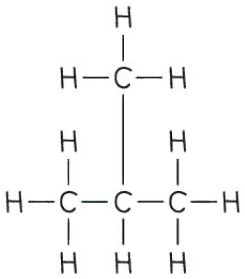
Structural Isomers

- Structural isomers are compounds that have the **same molecular** formula but **different structural** formulae
- The **molecular formula** is the **actual number** of atoms of each element in a compound
- Compounds with the same molecular formula can have different structural formulae due to the different arrangement of their atoms in space
- Two examples of structural isomers are shown below

Table showing Structural Isomerism in C_4H_{10}

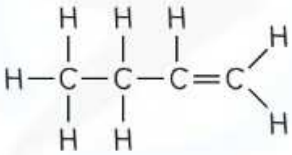
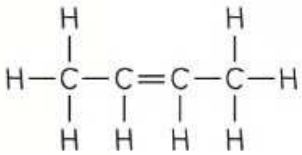
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Your notes

Molecular formula	C_4H_{10}	C_4H_{10}
Structural formula	$CH_3CH_2CH_2CH_3$	$CH_3CH(CH_3)CH_3$
Displayed formula		

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Table showing Structural Isomerism in C_4H_8

Molecular formula	C_4H_8	C_4H_8
Structural formula	$CH_3CH_2CH=CH_2$	$CH_3CH=CHCH_3$
Displayed formula		

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Exam Tip

Remember: Only double and triple bonds are shown in structural formulae.

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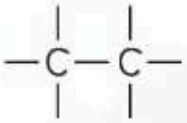
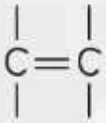
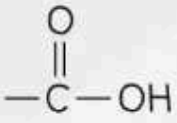
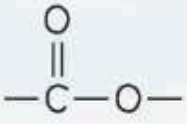
Homologous Series

Your notes

Homologous Series

- This is a series or family of organic compounds that have similar features and chemical properties due to them having the same **functional group**
- The functional group is a group of atoms which are bonded in a specific arrangement that is responsible for the characteristic reactions of each member of a homologous series

Table of Compounds & their Functional Groups

Family	Functional group	Name
Alkane		-ane
Alkene		-ene
Alcohol	—OH	-anol
Carboxylic acid		-anoic acid
Amine	$-NH_2$	-amine
Ester		-yl-anoate

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Exam Tip

Make sure you can identify the functional group for each homologous series.

Your notes

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General Formulae

Your notes

- This type of formula tells you the composition of any member of a whole homologous series of organic compound
- For example, all of the alkanes have the general formula C_nH_{2n+2}, where n represents the number of carbon atoms
- This tells you that however many carbon atoms there are in the alkane, doubling this number and adding two will give you the number of hydrogen atoms present in the alkane
- General formulae can be used to work out the formula of a compound from different homologous series if the number of carbon atoms present is known

General Formula of Common Homologous Series

Homologous Series	General Formula
Alkanes	C_nH_{2n+2}
Alkenes	C_nH_{2n}
Alcohols	$C_nH_{2n+1}OH$
Carboxylic Acids	$C_nH_{2n+1}COOH$

Worked example

What is the formula of an alcohol that contains 5 carbon atoms?

Answer

- Number of carbons = 5
- Number of hydrogens (excluding in the functional group) = 2 x 5 + 1 = 11
- Formula = $C_5H_{11}OH$

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Your notes

Worked example

A compound has the formula $C_{12}H_{24}$. To which homologous series does this compound belong to?

Answer

- There are 12 carbon atoms, so n = 12
- There are twice the number of hydrogen atoms than carbon atoms = 2n
- Therefore the general formula of the compound is C_nH_{2n} which means this compound is an alkene

General Characteristic of Homologous Series

- All members of a homologous series have:
 - The same **general formula**
 - Same **functional group**
 - Similar **chemical properties**
 - Gradation in their physical properties, such as melting and boiling point
 - The difference in the molecular formula between one member and the next is CH_2
 - These characteristics are shown below for ethanol and propanol, which belong to homologous series, alcohols

Table of Characteristics of Ethanol and Propanol

	Ethanol	Propanol
Molecular formula	CH_3CH_2OH	$CH_3CH_2CH_2OH$
General formula	$C_nH_{2n+1}OH$	$C_nH_{2n+1}OH$
Functional group	−OH	−OH
Boiling point (oC)	78	98

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Exam Tip

Make sure you learn the general formula for each homologous series.

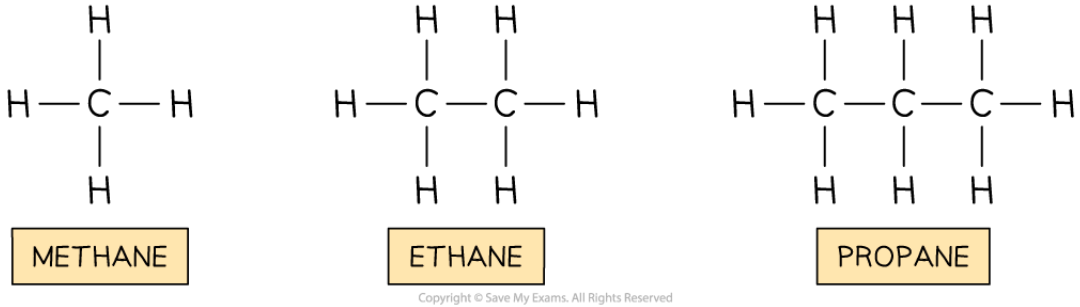
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Saturated & Unsaturated Compounds

Your notes

Saturated & Unsaturated Compounds

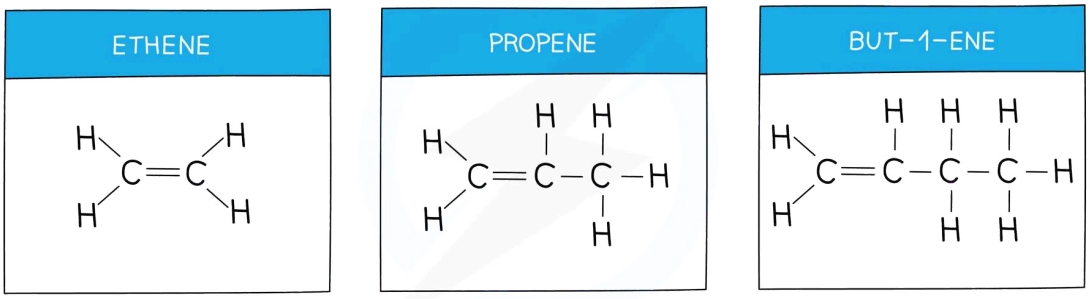
- **Saturated compounds** have molecules in which all carbon-carbon bonds are **single bonds**
- Examples of compounds that are saturated are **alkanes**
- Alkanes are saturated hydrocarbons with the general formula $\mathbf{C_nH_{2n+2}}$



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Alkanes contain only carbon-carbon single bonds so are saturated

- **Unsaturated compounds** consist of molecules in which one or more carbon-carbon bonds are **not** single bonds
- They contain carbon-carbon **double bonds** (C=C)
- Examples of compounds that are unsaturated are alkenes
- **Alkenes** are unsaturated hydrocarbons with the general formula is $\mathbf{C_nH_{2n}}$
- The presence of the double bond, C=C, means they can **make more bonds** with other atoms by opening up the C=C bond and allowing incoming atoms to form another single bond with each carbon atom of the functional group
- Each of these carbon atoms now forms 4 single bonds instead of 1 double and 2 single bonds



Alkenes contain one carbon-carbon double bond so are unsaturated

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Exam Tip

Remember: **S**aturated compounds have **S**ingle bonds only. **U**nsaturated compounds have do**U**ble bonds

Your notes

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Cambridge O Level Chemistry

Your notes

11.2 Naming Organic Compounds

Contents

Naming Organic Compounds

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Naming Organic Compounds

Your notes

Naming Organic Compounds

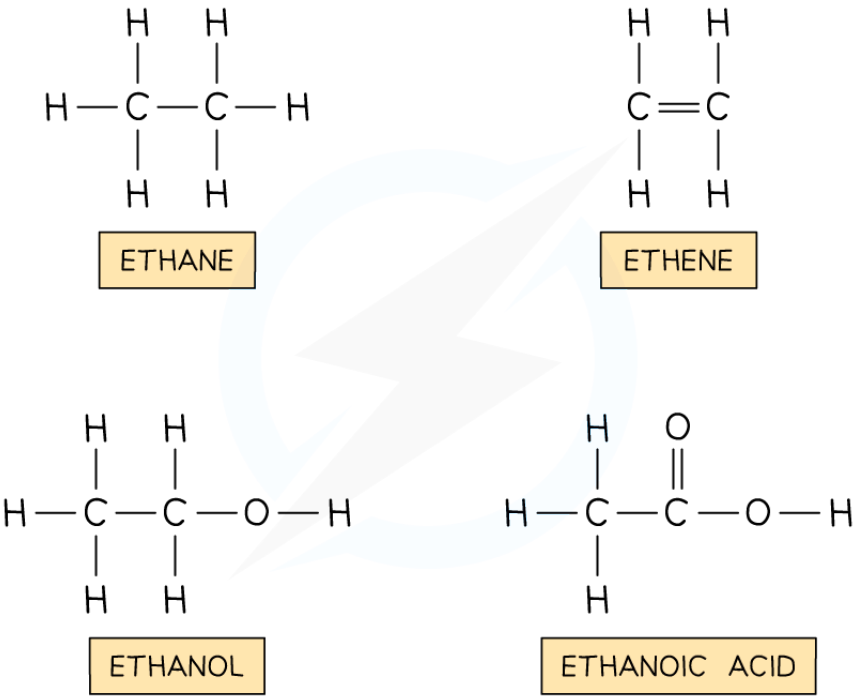
- The names of organic compounds have two parts: the prefix (or stem) and the end part (or suffix)
- The prefix tells you how many carbon atoms are present in the longest continuous chain in the compound
- The suffix tells you what functional group is on the compound

FIRST PART OF NAME		SECOND PART OF NAME		
NAME	NUMBER OF CARBON ATOMS	NAME	FUNCTIONAL GROUP	FAMILY
METH...	1	...ANE	NONE	ALKANE
ETH...	2	...ENE	C=C BOND	ALKENE
PROP...	3	...ANOL	R–OH	ALCOHOL
BUT...	4	...ANOIC ACID	R–C=O–OH	CARBOXYLIC ACID
PENT...	5	...AMINE	$R–NH_2$	AMINE
HEX...	6	...YL ...ANOATE	R–C=O–O–R	ESTER

Structures of organic compounds

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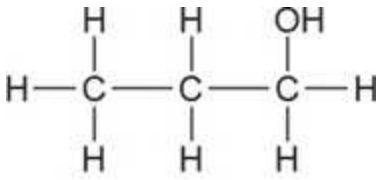
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Further Rules for Naming Compounds

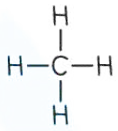
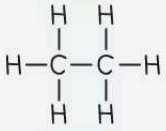
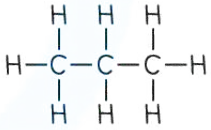
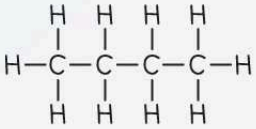
- When there is more than one carbon atom where a functional group can be located it is important to distinguish exactly **which carbon** the functional group is on
- Each carbon is numbered and these numbers are used to describe where the functional group is
- For example:
 - Propan-1-ol is alcohol with an -OH functional group
 - The 2 in the name indicates that the -OH group is located on the second carbon atom



In propan-1-ol the -OH group is located on the first carbon atom

Alkanes

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Alkane	Structural formula	Displayed formula
Methane	CH_4	
Ethane	CH_3CH_3	
Propane	$CH_3CH_2CH_3$	
Butane	$CH_3CH_2CH_2CH_3$	

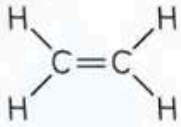
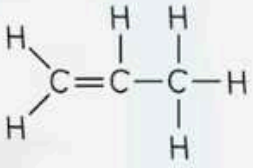
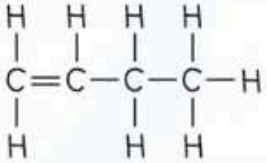
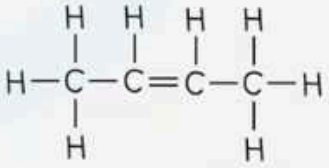
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Alkenes

Your notes

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Your notes

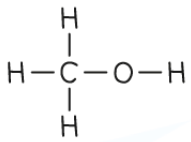
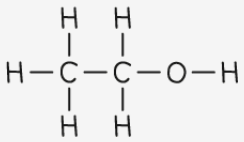
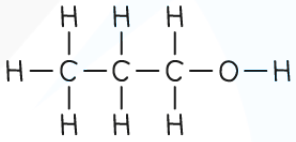
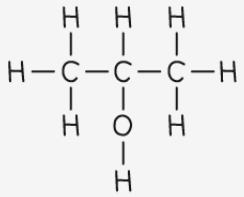
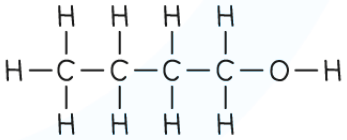
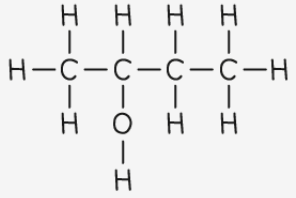
Alkene	Structural formula	Displayed formula
Ethene	$CH_2=CH_2$	
Propene	$CH_2=CHCH_3$	
But-1-ene	$CH_2=CHCH_2CH_3$	
But-2-ene	$CH_3CH=CHCH_3$	

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Alcohols

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Your notes

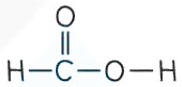
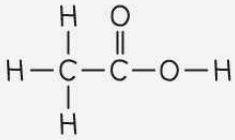
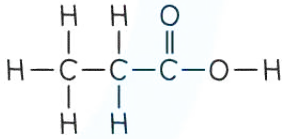
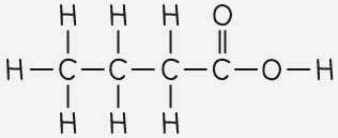
Alcohol	Structural formula	Displayed formula
Methanol	CH_3OH	
Ethanol	CH_3CH_2OH	
Propan-1-ol	$CH_3CH_2CH_2OH$	
Propan-2-ol	$CH_3CHOHCH_3$	
Butan-1-ol	$CH_3CH_2CH_2CH_2OH$	
Butan-2-ol	$CH_3CH_2CHOHCH_3$	

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Your notes

Carboxylic acids

Carboxylic acid	Structural formula	Displayed formula
Methanoic acid	HCOOH	
Ethanoic acid	CH_3COOH	
Propanoic acid	CH_3CH_2COOH	
Butanoic acid	$CH_3CH_2CH_2COOH$	

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Exam Tip

Make sure you can draw and name the structures given above.

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Naming Esters

Your notes

- An ester is made from an alcohol and carboxylic acid
- The first part of the name indicates the length of the carbon chain in the alcohol, and it ends with the letters '- yl'
- The second part of the name indicates the length of the carbon chain in the carboxylic acid, and it ends with the letters '- oate'
- e.g. the ester formed from **pent**anol and **butan**oic acid is called **pent**yl **butan**oate

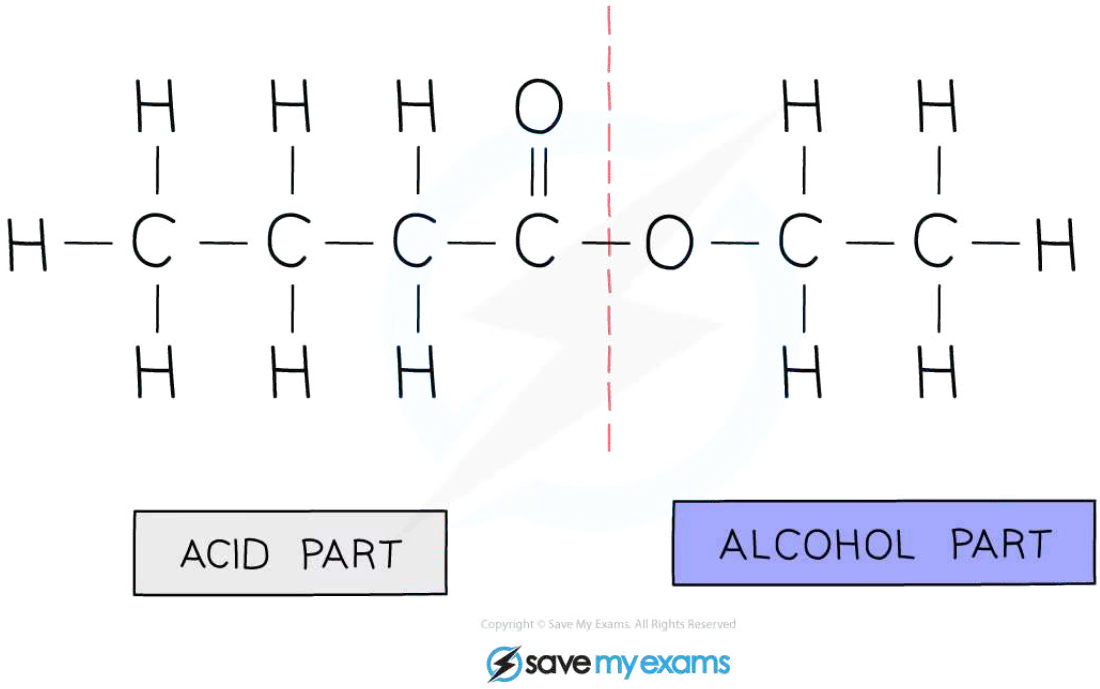


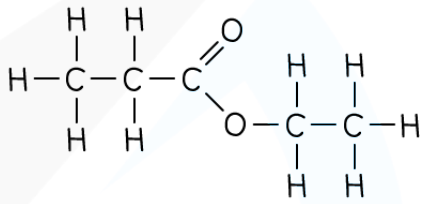
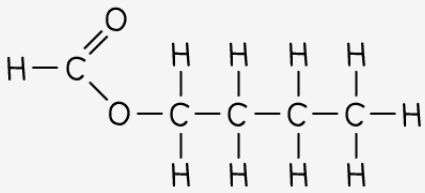
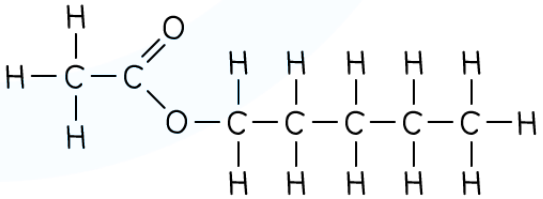
Diagram showing the origin of each carbon chain in ester

- Some examples of common esters:

Table showing the Formation of Esters

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Your notes

NAME OF ALCOHOL	NAME OF CARBOXYLIC ACID	NAME OF ESTER
ETHANOL	PROPANOIC ACID	ETHYL PROPANOATE 
BUTANOL	METHANOIC ACID	BUTYL METHANOATE 
PENTANOL	ETHANOIC ACID	PENTYLETHANOATE 

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Exam Tip

You need to be able to draw and name esters made from alcohols and carboxylic acids each containing up to four carbon atoms.

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Cambridge O Level Chemistry

Your notes

11.3. Fuels

Contents

Fossil Fuels

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Fossil Fuels


Your notes

Common Fossil Fuels

- A fuel is a substance which when burned, releases heat energy
- This heat can be transferred into electricity, which we use in our daily lives
- Most common fossil fuels include coal, natural gas and **hydrocarbons** such as methane and propane which are obtained from crude oil
- Hydrocarbons are made from hydrogen and carbon atoms only
- The main constituent of natural gas is **methane**, CH_4

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Petroleum & Fractional Distillation

Petroleum

Your notes

- Petroleum is also called **crude oil** and is a complex mixture of hydrocarbons which also contains natural gas
- It is a thick, sticky, black liquid that is found under porous rock (under the ground and under the sea)

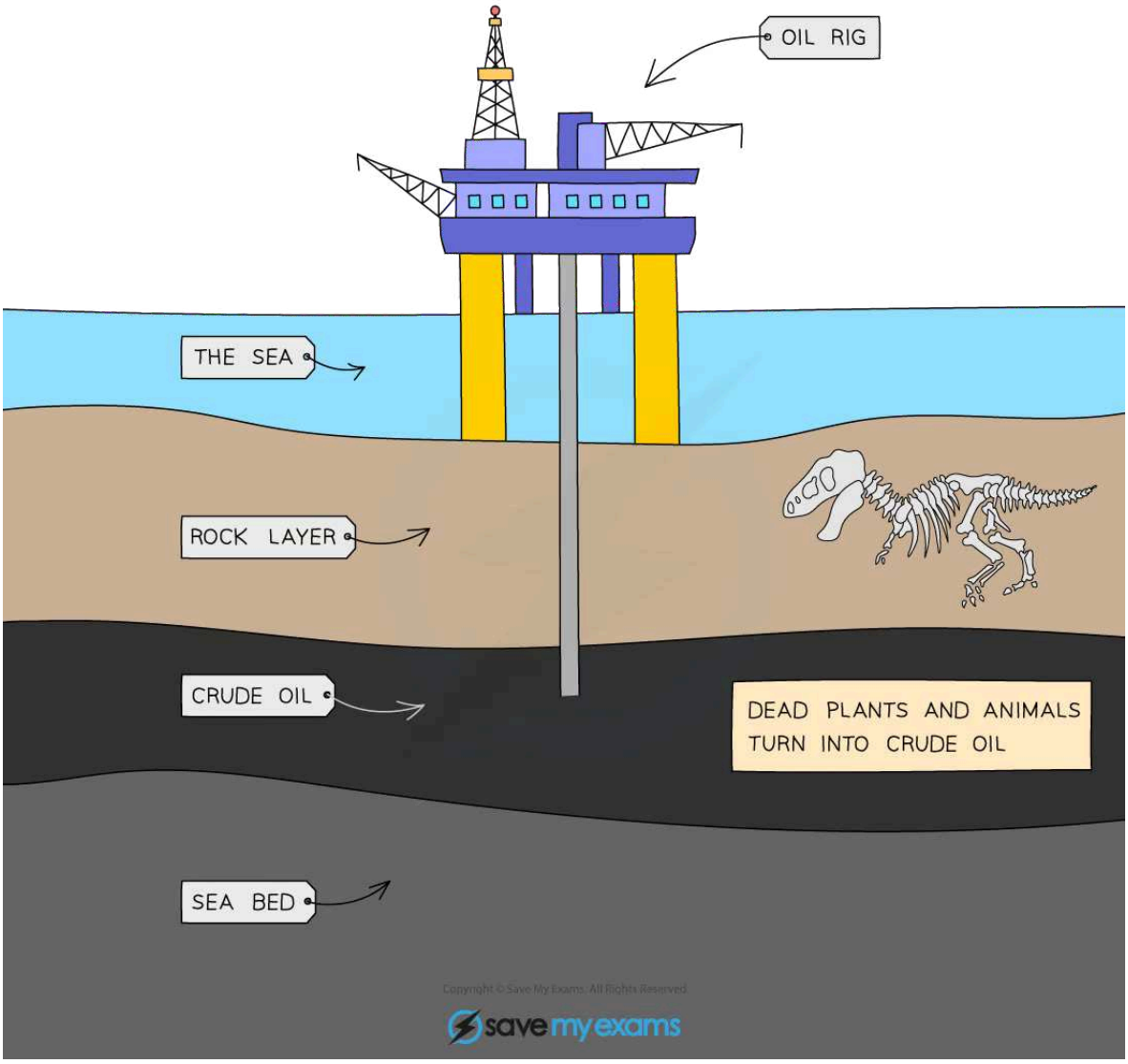


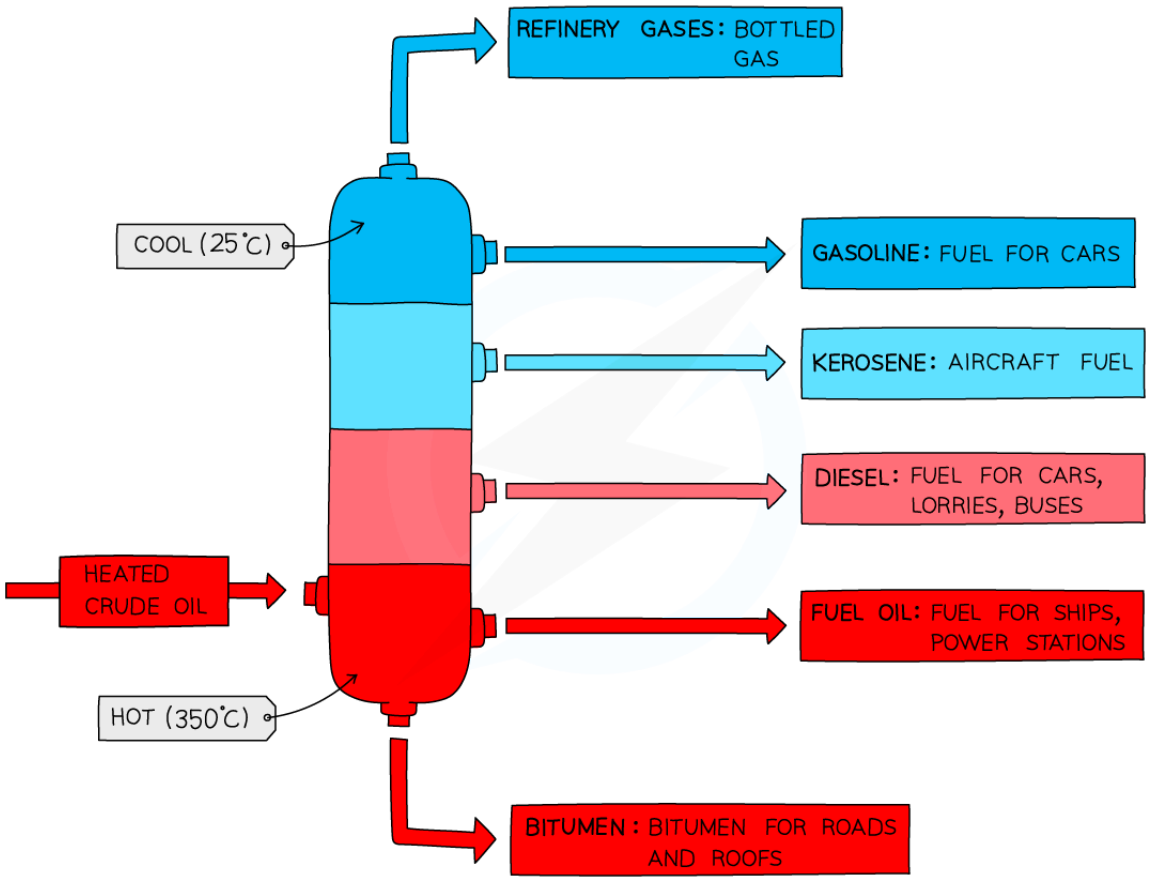
Diagram showing crude oil under the sea

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- Petroleum itself as a mixture isn't very useful but each component part of the mixture, called a fraction, is useful and each fraction has different applications
- The fractions in petroleum are separated from each other in a process called **fractional distillation**
- The molecules in each fraction have similar **properties** and **boiling points**, which depend on the number of carbon atoms in the chain
- The **boiling point** and viscosity of each fraction increase as the carbon chain gets longer

Your notes

Fractional Distillation



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Diagram showing the process of fractional distillation

- Fractional distillation is carried out in a fractionating column
- The fractionating column is hot at the bottom and cools at the top
- Crude oil enters the fractionating column and is **heated** so **vapours** rise

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- Vapours of hydrocarbons with very high boiling points will immediately turn into liquid and are tapped off at the bottom of the column
- Vapours of hydrocarbons with low boiling points will rise up the column and condense at the top to be tapped off
- The different fractions condense at different heights according to their boiling points and are tapped off as liquids.
- The fractions containing smaller hydrocarbons are collected at the top of the fractionating column as gases
- The fractions containing bigger hydrocarbons are collected at the lower sections of the fractionating column


Your notes

Properties of Fractions

- **Viscosity**
 - This refers to the ease of flow of a liquid.
 - High viscosity liquids are thick and flow less easily.
 - If the number of carbon atoms increases, the attraction between the hydrocarbon molecules also increases which results in the liquid becoming more viscous with the increasing length of the hydrocarbon chain.
 - The liquid flows less easily with increasing molecular mass
- **Colour**
 - As carbon chain length increases the colour of the liquid gets darker as it gets thicker and more viscous
- **Melting point/boiling point**
 - As the molecules get larger, the intermolecular attraction becomes greater.
 - More heat is needed to separate the molecules.
 - With increasing molecular size there is an increase in boiling point
- **Volatility**
 - Volatility refers to the tendency of a substance to vaporise.
 - With increasing molecular size hydrocarbon liquids become less volatile.
 - This is because the attraction between the molecules increases with increasing molecular size

Uses of Fractions

- Refinery gas: heating and cooking
- Gasoline: fuel for cars (petrol)
- Naphtha: raw product for producing chemicals
- Kerosene: for making jet fuel (paraffin)
- Diesel: fuel for diesel engines (gas oil)
- Fuel oil: fuel for ships and for home heating
- Lubricating oil: for lubricants, polishes, waxes

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- Bitumen: for surfacing roads

Trends in Properties

FRACTION	NUMBER OF CARBON ATOMS	BOILING POINT RANGE /°C	BOILING POINT & VISCOSITY INCREASE GOING DOWN
REFINERY GAS	1 – 4	BELOW 25	
GASOLINE / PETROL	4 – 12	40 – 100	
NAPHTHA	7 – 14	90 – 150	
KEROSENE / PARAFFIN	12 – 16	150 – 240	
DIESEL / GAS OIL	14 – 18	220 – 300	
FUEL OIL	19 – 25	250 – 320	
LUBRICATING OIL	20 – 40	300 – 350	
BITUMEN	MORE THAN 70	MORE THAN 350	

Exam Tip

When defining a hydrocarbon, ensure you say that it has hydrogen and carbon atoms **only.**

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Cambridge O Level Chemistry

Your notes

11.4 Alkanes

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Alkanes

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Alkanes


Your notes

Alkanes: Properties & Bonding

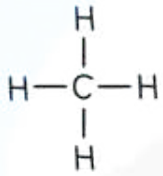
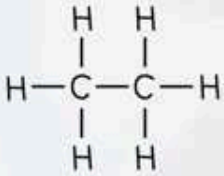
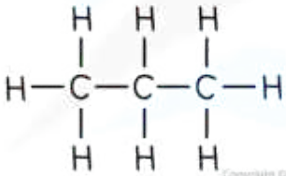
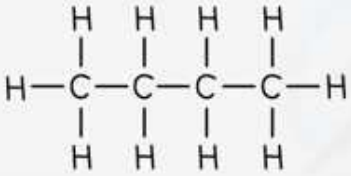
- Alkanes are a group of **saturated** hydrocarbons
- The term saturated means that they only have single carbon-carbon bonds, there are no double bonds
- The general formula of the alkanes is $\mathbf{C_nH_{2n+2}}$
- They are colourless compounds which have a gradual change in their physical properties as the number of carbon atoms in the chain increases
- Alkanes are generally unreactive compounds but they do undergo **combustion** reactions, can be **cracked** into smaller molecules and can react with **halogens** in the presence of light in substitution reactions
- Methane is an alkane and is the major component of **natural gas**
- Methane undergoes complete combustion forming carbon dioxide and water:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

This Table shows the Displayed Formula of the First Four Members of the Alkane Homologous Series

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Your notes

Structural Formula	Name	Molecular Formula
	methane	CH_4
	ethane	C_2H_6
	propane	C_3H_8
	butane	C_4H_{10}

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Substitution Reaction of Alkanes with Halogens

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- In a **substitution reaction**, one atom (or group of atoms) is swapped with another atom (or group of atoms)
- Alkanes undergo a substitution reaction with halogens in the presence of ultraviolet radiation (sunlight is a source of UV radiation)
- This is called a **photochemical** reaction
- The UV light provides the **activation energy**, E_a, for the reaction
- A hydrogen atom is replaced with the halogen atom
- More than one hydrogen atom can be substituted depending on the amount of ultraviolet radiation there is



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In the presence of ultraviolet (UV) radiation, methane reacts with chlorine to form chloromethane and hydrogen chloride

Exam Tip

You need to be able to draw the displayed and structural formulae of the products formed when one halogen atom replaces a hydrogen (also known as monosubstitution)

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11.5 Alkenes

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- Alkenes
- Addition Reactions

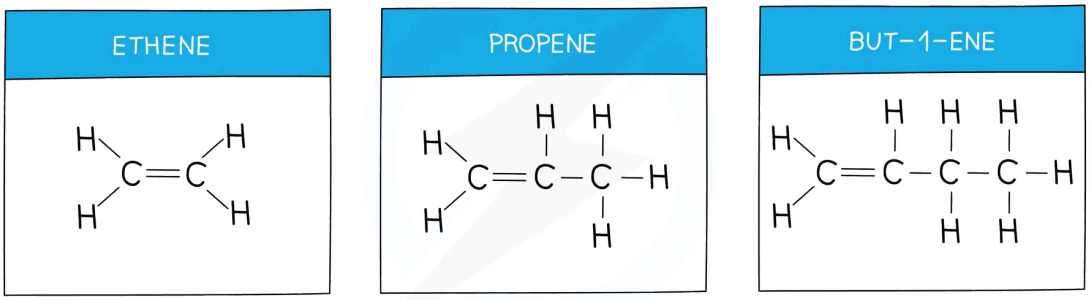
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Alkenes

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Catalytic Cracking

- Alkenes are unsaturated hydrocarbons with carbon-carbon double bonds (C=C)
- Their general formula is C_nH_{2n}
- The presence of the double bond, C=C, means they can **make more bonds** with other atoms by opening up the C=C bond and allowing incoming atoms to form another single bond with each carbon atom of the functional group
- Each of these carbon atoms now forms 4 single bonds instead of 1 double and 2 single bonds
- This makes them much more reactive than alkanes



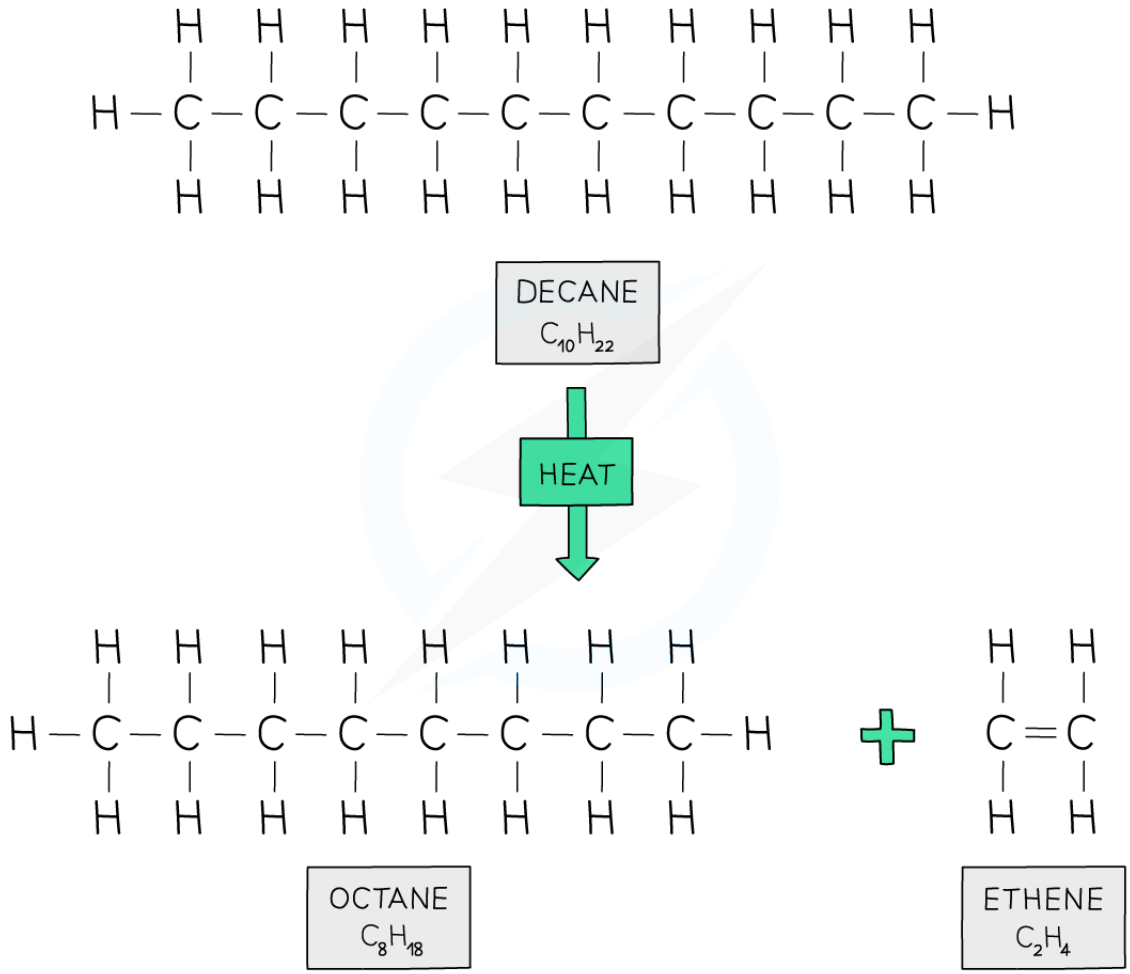
The displayed formula of the first three alkenes

Manufacture of Alkenes

- Although there is use for each fraction obtained from the fractional distillation of crude oil, the amount of longer chain hydrocarbons produced is far greater than needed
- These long chain hydrocarbon molecules are further processed to produce other products
- A process called catalytic cracking is used to convert longer-chain molecules into short-chain and more useful hydrocarbons
- Shorter chain alkanes, alkenes and hydrogen are produced from the cracking of longer chain alkanes
- Alkenes can be used to make polymers and the hydrogen used to make ammonia
- Kerosene and diesel oil are often cracked to produce petrol, other alkenes and hydrogen
- **Cracking** involves heating the hydrocarbon molecules to around 600 - 700°C to **vaporise** them
- The vapours then pass over a hot powdered **catalyst** of alumina or silica
- This process breaks covalent bonds in the molecules as they come into contact with the surface of the catalyst, causing **thermal decomposition** reactions
- The molecules are broken up in a random way which produces a mixture of smaller alkanes and alkenes
- Hydrogen and a higher proportion of alkenes are formed at higher temperatures and higher pressure

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The 10 carbon molecule decane is catalytically cracked to produce octane for petrol and ethene for ethanol

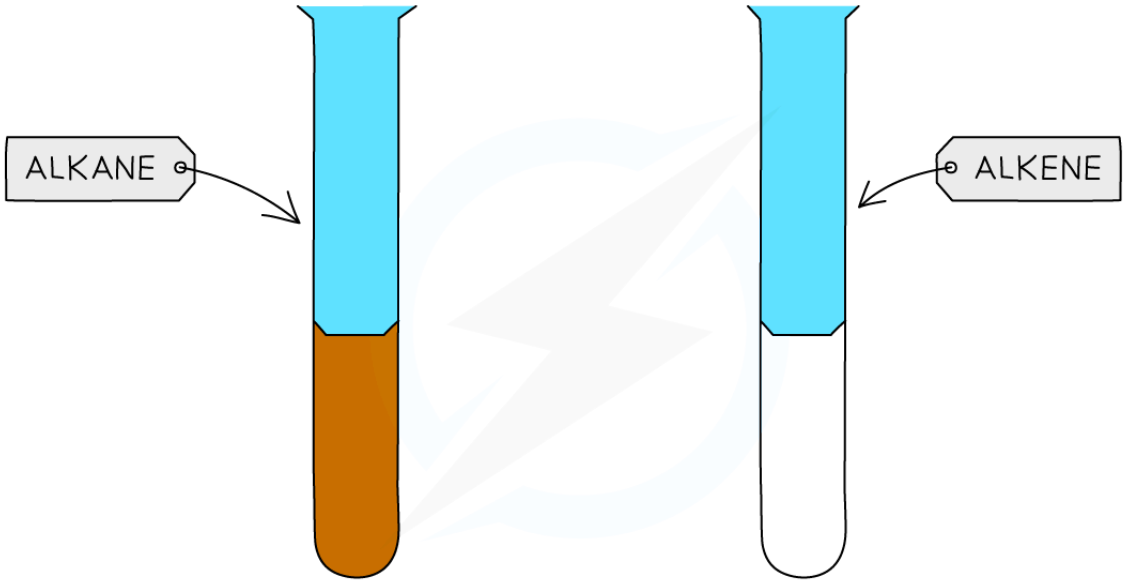
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Distinguishing from Alkanes

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Distinguishing Between Alkanes & Alkenes

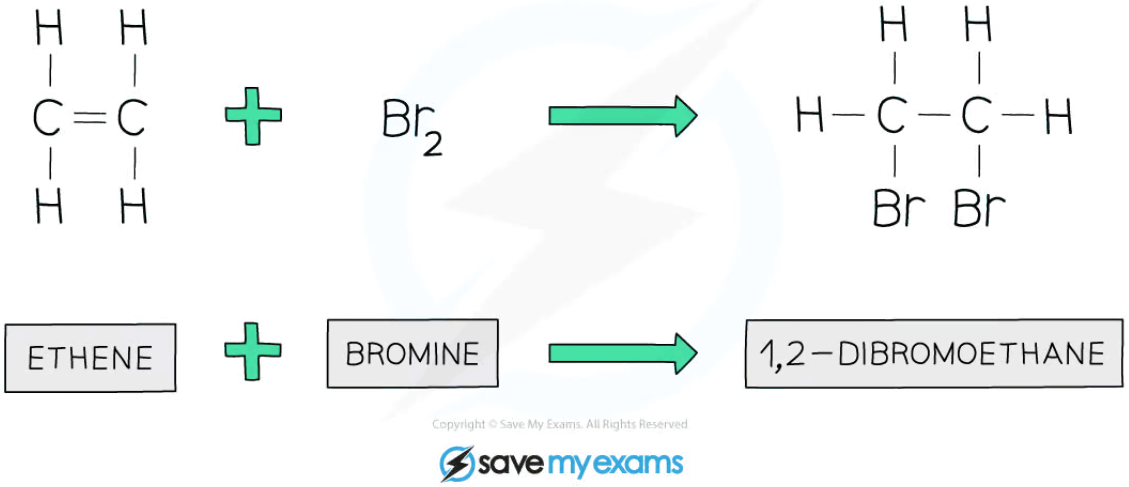
- Alkanes and alkenes have different molecular structures
- All alkanes are saturated and alkenes are unsaturated
- The presence of the C=C double bond allows alkenes to react in ways that alkanes cannot
- This allows us to tell alkenes apart from alkanes using a simple chemical test using **bromine water**



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- Bromine water is an orange coloured solution of bromine
- When bromine water is shaken with an alkane, it will remain as an orange solution as alkanes do not have double carbon bonds (C=C) so the bromine remains in solution
- When bromine water is shaken with an alkene, the alkene will decolourise the bromine water and turn colourless as alkenes do have double carbon bonds (C=C)
- The bromine atoms add across the C=C double bond hence the solution no longer contains the orange coloured bromine
- This reaction between alkenes and bromine is called an addition reaction

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Each carbon atom of the double bond accepts a bromine atom, causing the bromine solution to lose its colour

Exam Tip

When describing what happens to bromine water in an alkene ensure you say colourless, and not clear.

Your notes

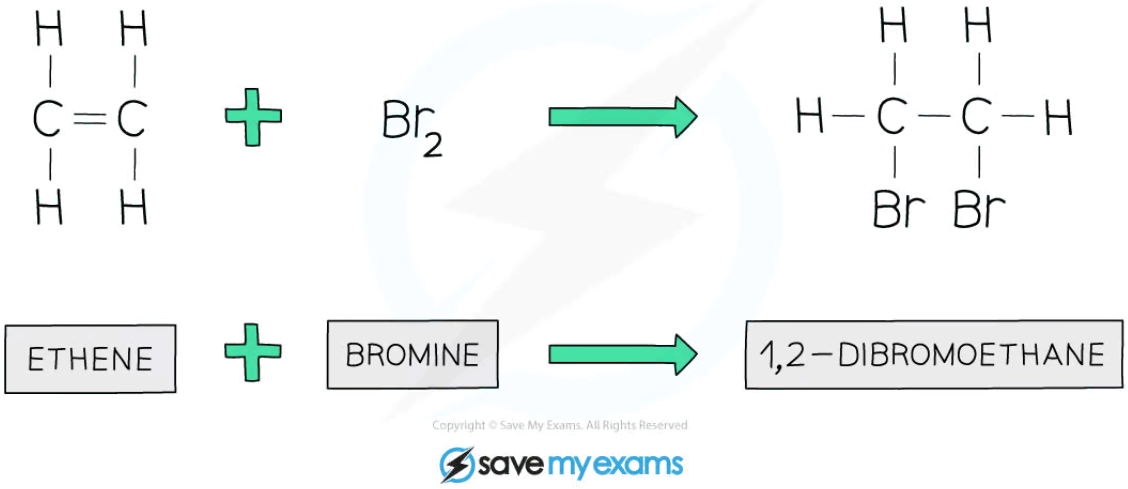
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Addition Reactions

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Addition Reactions

- Alkenes undergo addition reactions in which atoms of a simple molecule add across the C=C double bond
- The reaction between bromine and ethene is an example of an addition reaction



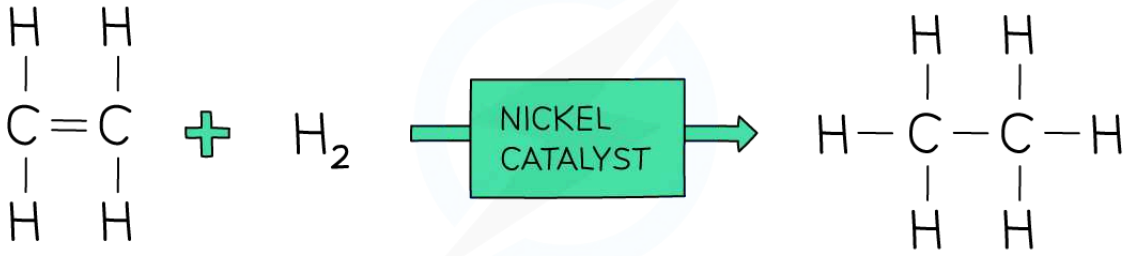
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Bromine atoms add across the C=C in the addition reaction of ethene and bromine

- Alkenes also undergo addition reactions with hydrogen in which an **alkane** is formed
- These are **hydrogenation reactions** and occur at **150°C** using a **nickel catalyst**
- Hydrogenation reactions are used to manufacture margarine from vegetable oils
 - Vegetable oils are polyunsaturated molecules which are partially hydrogenated to increase the M_r and turn the oils into solid fats

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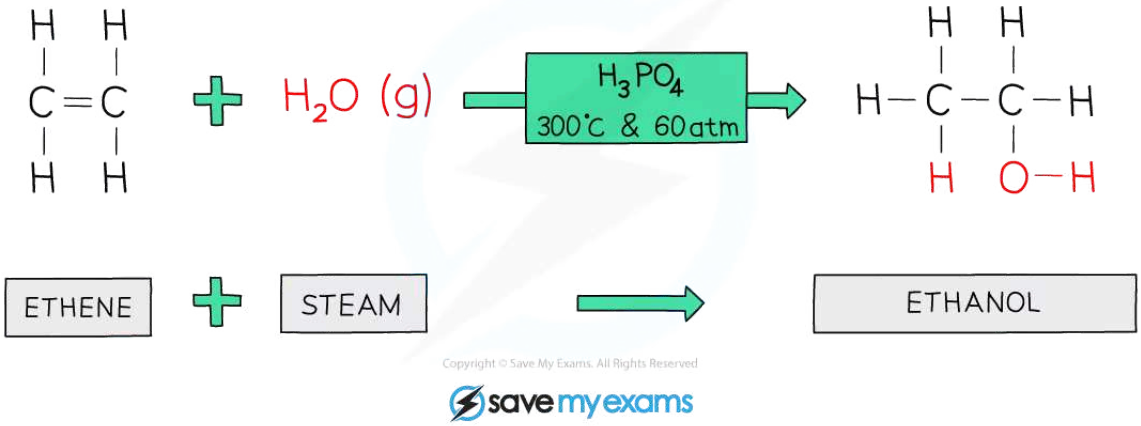
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Hydrogen atoms add across the C=C in the hydrogenation of ethene to produce an alkane

- Alkenes also undergo addition reactions with steam in which an **alcohol** is formed.
- Since water is being added to the molecule it is also called a **hydration** reaction
- The reaction is very important industrially for the production of alcohols and it occurs using the following conditions:
 - Temperature of around **300°C**
 - Pressure of **60 - 70 atm**
 - Concentrated phosphoric **acid catalyst**



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A water molecule adds across the C=C in the hydration of ethene to produce ethanol

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Exam Tip

You need to be able to draw the displayed formulae of the products of alkenes with water, hydrogen and bromine.

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Cambridge O Level Chemistry

Your notes

11.6 Alcohols

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Alcohols

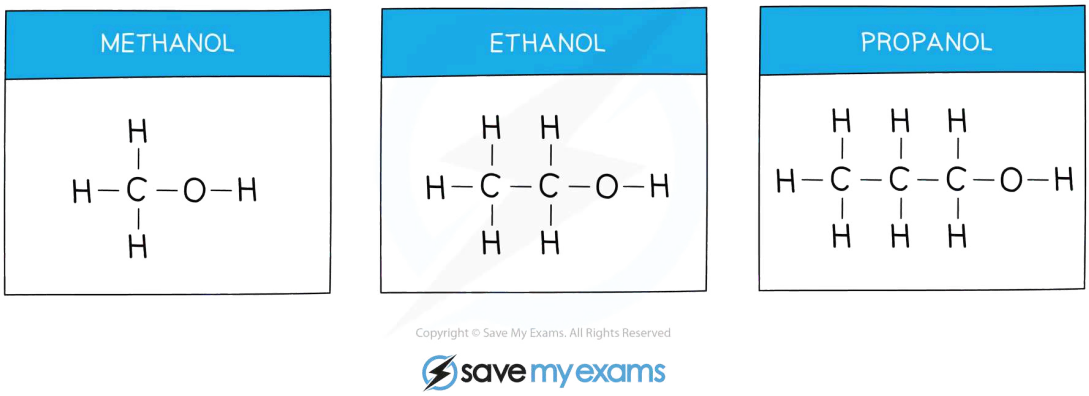
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Alcohols

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Alcohols

- All alcohols contain the hydroxyl (**-OH**) functional group which is the part of alcohol molecules that is responsible for their **characteristic reactions**
- Alcohols are a homologous series of compounds that have the general formula $\mathbf{C_nH_{2n+1}OH}$
- They differ by one -CH_2 in the molecular formulae from one member to the next



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Diagram showing the first three alcohols

- Ethanol (C_2H_5OH) is one of the most important alcohols
 - Ethanol can also be represented by its structural formula CH_3CH_2OH
- It is the type of alcohol found in **alcoholic drinks** such as wine and beer
- It is also used as fuel for **cars** and as a **solvent**
- Alcohols burn in excess oxygen and produce CO_2 and H_2O
- Ethanol undergoes complete combustion:

$$C_2H_5OH\,(l) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 3H_20\,(l)$$

The Manufacture of Ethanol

- There are two methods used to manufacture ethanol:
 - The **hydration** of **ethene** with steam
 - The **fermentation** of glucose
- Both methods have advantages and disadvantages which are considered

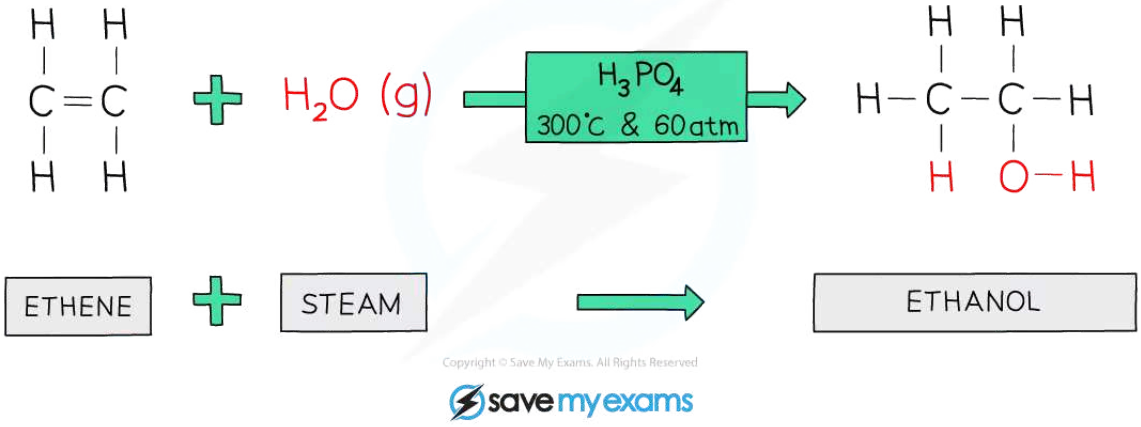
Hydration of ethene

- A mixture of ethene and steam is passed over a hot catalyst of **phosphoric acid** at a temperature of approximately **300 °C**
- The pressure used is **60 atmospheres (6000 kPa)**

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- The gaseous ethanol is then condensed into a liquid for use

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A water molecule adds across the C=C in the hydration of ethene to produce ethanol

Fermentation of glucose

- Sugar or starch is dissolved in water and yeast is added
- The mixture is then fermented between **25** and **35 °C** with the **absence** of oxygen for a few days
- Yeast contains enzymes that catalyse the break down of starch or sugar to glucose
- If the temperature is too **low** the reaction rate will be too slow and if it is too **high** the enzymes will become denatured
- The yeast respire anaerobically using the glucose to form ethanol and carbon dioxide:

$$C_6H_{12}O_6 \rightarrow 2CO_2 + 2C_2H_5OH$$

- The yeast are killed off once the concentration of alcohol reaches around 15%, hence the reaction vessel is emptied and the process is started again
- This is the reason that ethanol production by fermentation is a batch process

Exam Tip

Make sure you learn the conditions for both hydration and fermentation.

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11.7 Carboxylic Acids

Contents

- Carboxylic Acids
- Ethanoic Acid & Esterification Reactions

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Carboxylic Acids


Your notes

Carboxylic Acids

- The carboxylic acids behave like other acids
- They react with:
 - **metals** to form a salt and hydrogen
 - **carbonates** to form a salt, water and carbon dioxide gas
- They also take part in **neutralisation** reactions to produce **salt** and **water**
- **Ethanoic acid** (also called **acetic acid**) is the acid used to make **vinegar**, which contains around 5% by volume of ethanoic acid
- The salts formed by the reaction of carboxylic acids all end –**anoate**
- So methanoic acid forms a salt called **methanoate**, **ethanoic** a salt called ethanoate etc.
- In the reaction with metals, a **metal** salt and hydrogen gas are produced

Example reactions of carboxylic acids

- The reaction of ethanoic acid with **magnesium** forms the salt magnesium **ethanoate** and hydrogen gas:

$$2CH_3COOH + Mg \rightarrow (CH_3COO)_2Mg + H_2$$

- In the reaction with **hydroxides,** salt and water are formed in a **neutralisation reaction**
- For example, the reaction between potassium hydroxide and propanoic acid forms the salt potassium propanoate and water:

$$CH_3CH_2COOH + KOH \rightarrow CH_3CH_2COOK + H_2O$$

- In the reaction with **carbonates** a metal salt, water and carbon dioxide gas are produced
- For example, in the reaction between potassium carbonate and butanoic acid, the salt potassium butanoate is formed with water and carbon dioxide

$$2CH_3CH_2CH_2COOH + K_2CO_3 \rightarrow 2CH_3CH_2CH_2COOK + H_2O + CO_2$$

Exam Tip

You need to be able to name and give the formulae of the salts produced in these reactions.

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Ethanoic Acid & Esterification Reactions


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Formation of Ethanoic Acid

Making Carboxylic Acids

- Two methods used to make carboxylic acids are:
 - Oxidation **by fermentation**
 - Using **oxidising agents**
- The **microbial oxidation** (fermentation) of ethanol will produce a weak solution of vinegar (ethanoic acid)
- This occurs when a bottle of wine is opened as bacteria in the air (acetobacter) will use atmospheric oxygen from air to oxidise the ethanol in the wine

$$C_2H_5OH\,(aq) + O_2\,(g) \rightarrow CH_3COOH\,(aq) + H_2O\,(l)$$

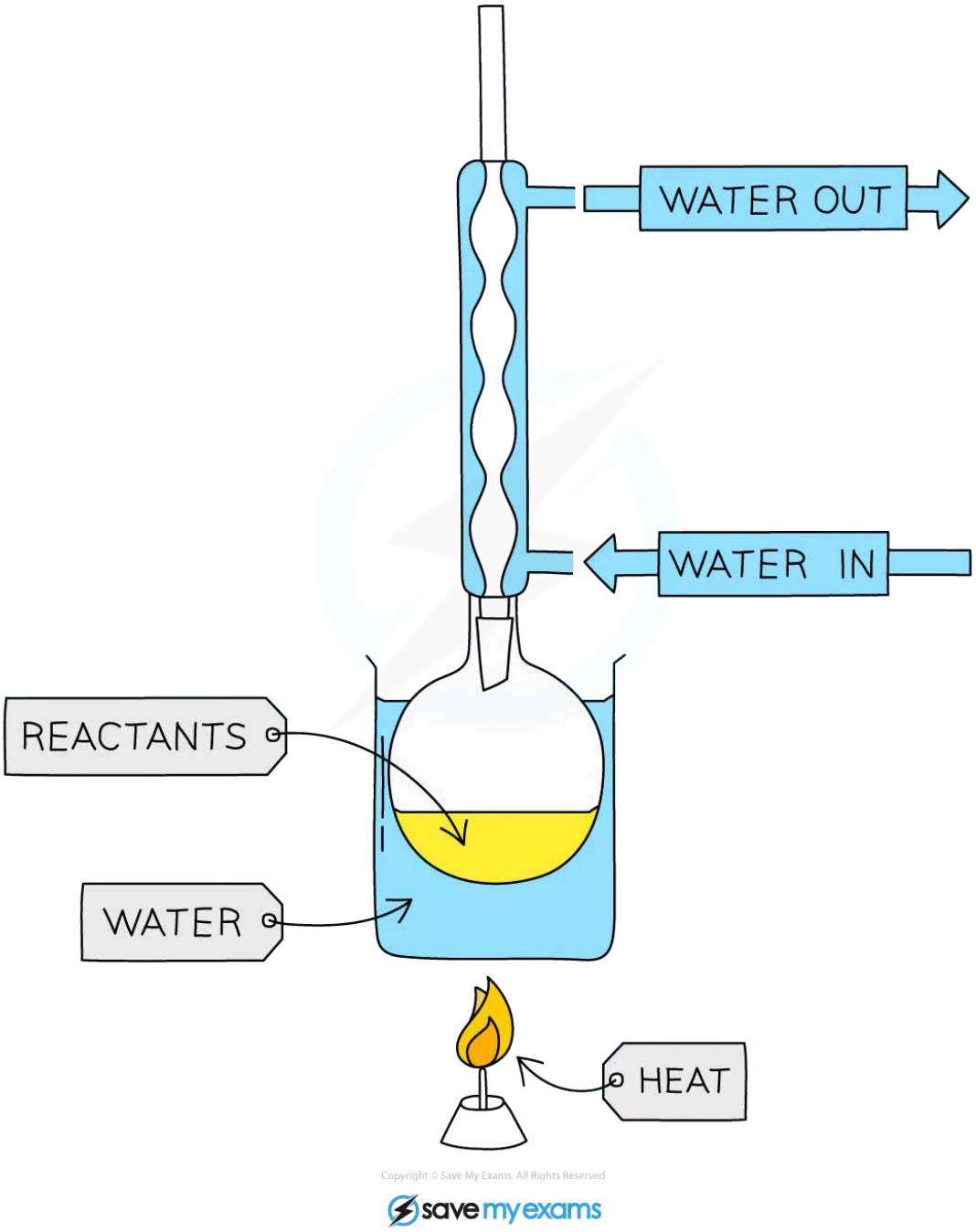
- The acidic, vinegary taste of wine which has been left open for several days is due to the presence of ethanoic acid
- Alternatively, oxidising agent **potassium manganate(VII)** can be used
- This involves heating ethanol with **acidified potassium manganate(VII)** in the presence of an acid
- The heating is performed under reflux which involves heating the reaction mixture in a vessel with a condenser attached to the top
- The condenser prevents the volatile alcohol from escaping the reaction vessel as alcohols have low boiling points
- The equation for the reaction is:

$$CH_3CH_2OH\,(aq) + [O] \rightarrow CH_3COOH\,(aq) + H_2O\,(l)$$

- The solution will change from purple to colourless
- The oxidising agent is represented by the symbol for oxygen in square brackets

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Diagram showing the experimental setup for the oxidation with $KMnO_4$ using reflux apparatus

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Esterification

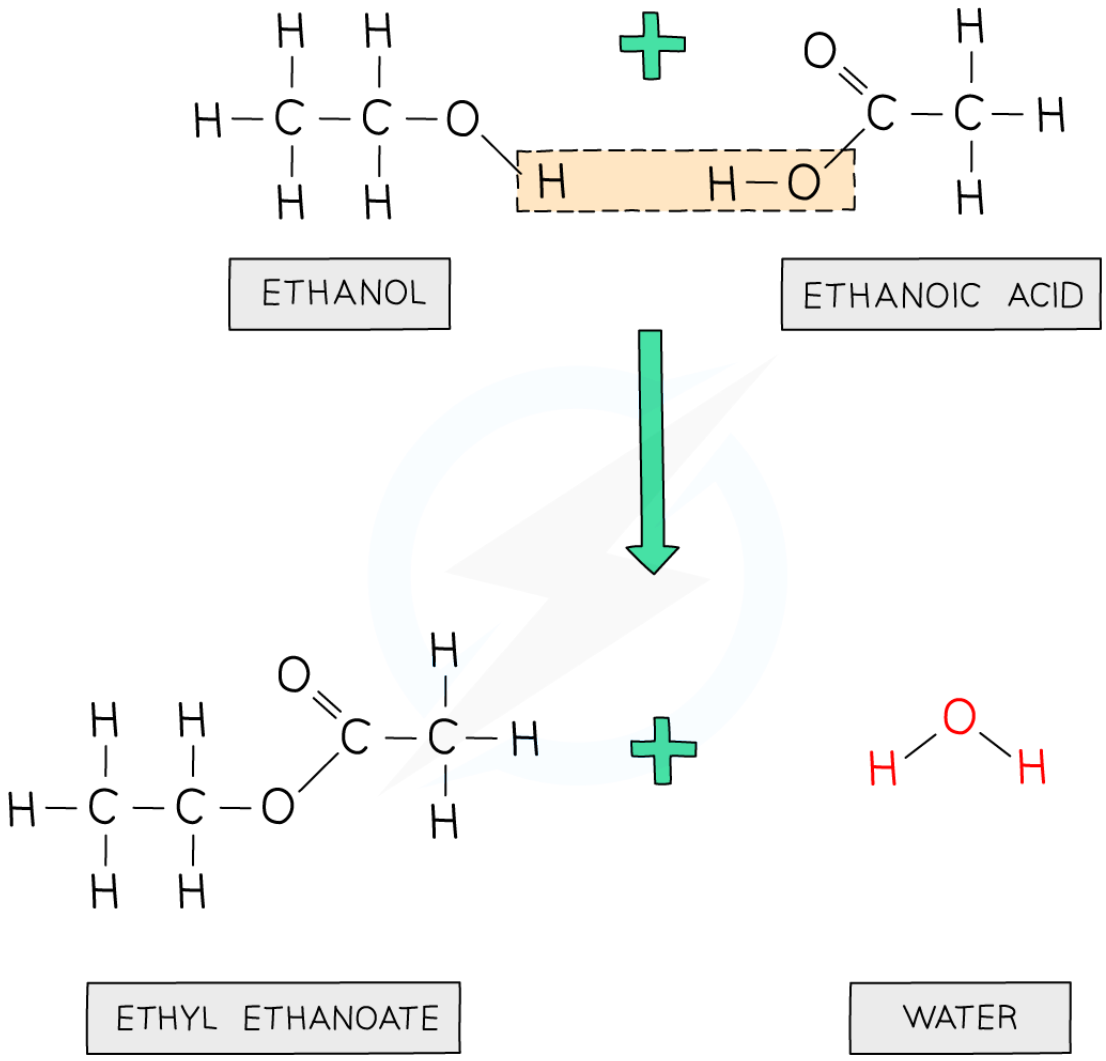

Your notes

- Alcohols and carboxylic acids react to make esters in **esterification** reactions
- Esters are compounds with the functional group R-COO-R
- Esters are sweet-smelling oily liquids used in food flavourings and perfumes
- Ethanoic acid will react with ethanol in the presence of concentrated sulfuric acid (catalyst) to form ethyl ethanoate:

$$CH_3COOH\,(aq) + C_2H_5OH\,(aq) \rightleftharpoons CH_3COOC_2H_5\,(aq) + H_2O\,(l)$$

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Diagram showing the formation of ethyl ethanoate

- For more information on how esters are named, see our revision note on Naming Esters

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11.8 Polymers

Contents

- Polymers
- Addition & Condensation Polymers
- Plastics & Their Disposal
- Proteins

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Polymers

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Polymers: The Basics

- Polymers are large molecules built by linking 50 or more smaller molecules called monomers
- Each repeat unit is connected to the adjacent units via **covalent bonds**
- Some polymers contain just one type of unit
 - Examples include poly(ethene) and poly(chloroethene), commonly known as PVC
- Others contain two or more different types of monomer units and which are called **copolymers**
 - Examples include nylon and biological proteins
- Different **linkages** also exist, depending on the monomers and the type of polymerisation
 - Examples of linkages are covalent bonds, amide links and ester links

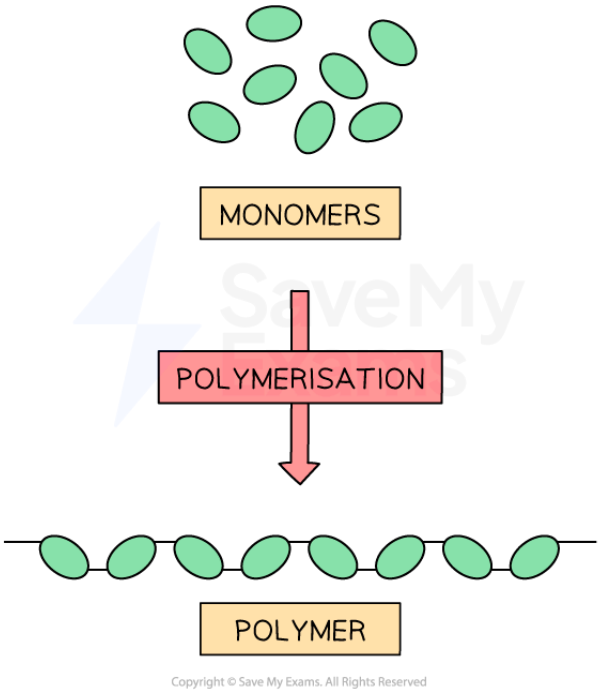
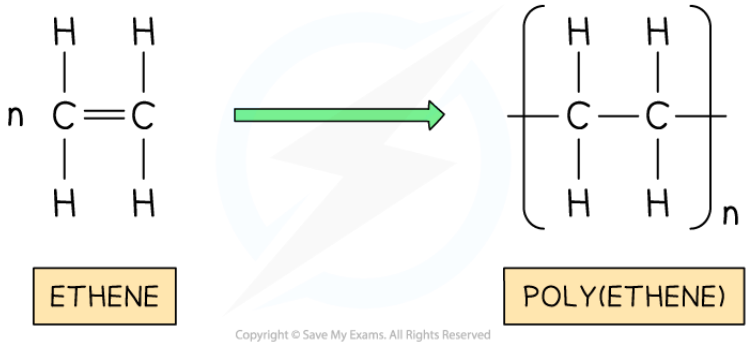


Diagram showing how lots of monomers bond together to form a polymer

- Poly(ethene) is formed by the addition polymerisation of ethene monomers
- Addition polymerisation involves the addition of many monomers to make a long chained polymer
- In this case, many ethene monomers join together due to the carbon carbon double bond breaking

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Poly(ethene) is formed by addition polymerisation using ethene monomers

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Addition & Condensation Polymers


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Addition Polymers

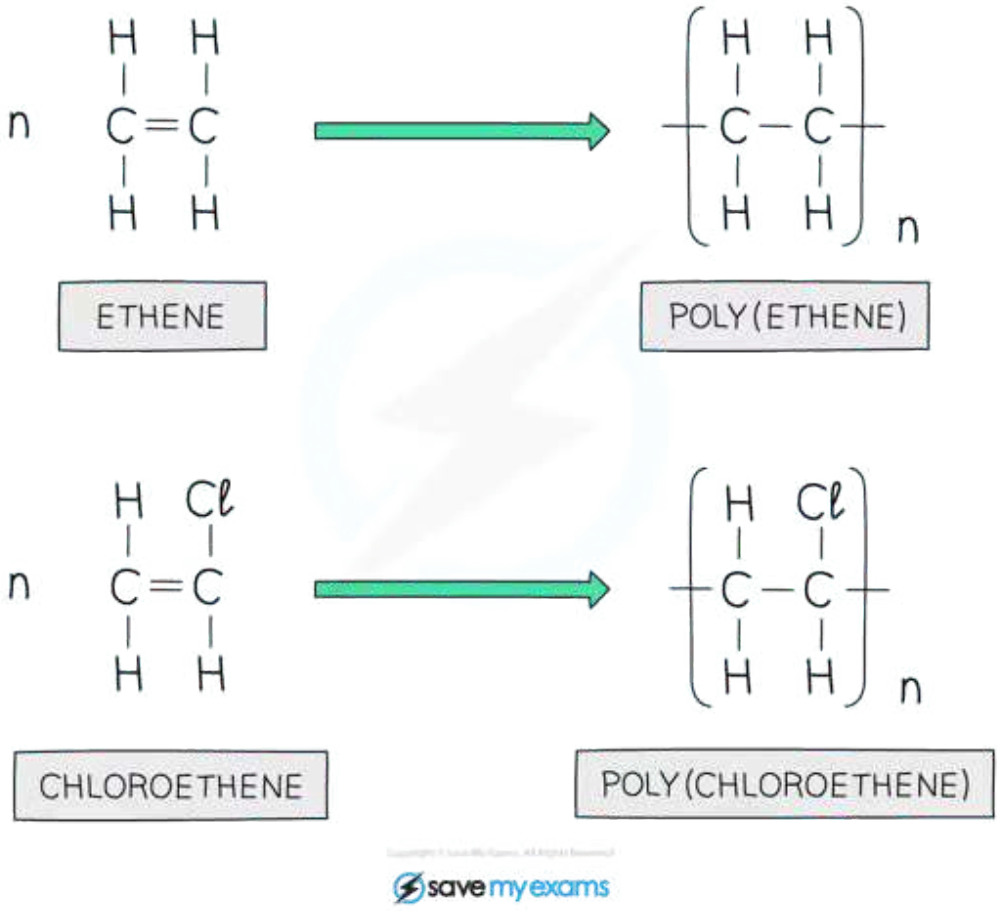
- Addition polymers are formed by the joining up of many monomers and only occur in monomers that contain **C=C bonds**
- **One** of the bonds in each C=C bond breaks and forms a bond with the adjacent monomer with the polymer being formed containing **single bonds** only
- Many polymers can be made by the addition of alkene monomers
- Others are made from alkene monomers with different atoms attached to the monomer such as chlorine or a hydroxyl group
- The name of the polymer is deduced by putting the name of the **monomer** in brackets and adding poly- as the **prefix**
- For example if propene is the alkene monomer used, then the name is **poly(propene)**
- Poly(ethene) is formed by the addition polymerisation of **ethene monomers**

Deducing the polymer from the monomer

- Polymer molecules are very large compared with most other molecule
- **Repeat units** are used when displaying the formula
- To draw a repeat unit, change the double bond in the monomer to a **single bond** in the repeat unit
- Add a bond to each end of the repeat unit
- The bonds on either side of the polymer must **extend** outside the brackets (these are called extension or continuation bonds)
- A small **subscript** n is written on the bottom right hand side to indicate a large number of repeat units
- Add on the rest of the groups in the **same order** that they surrounded the double bond in the monomer

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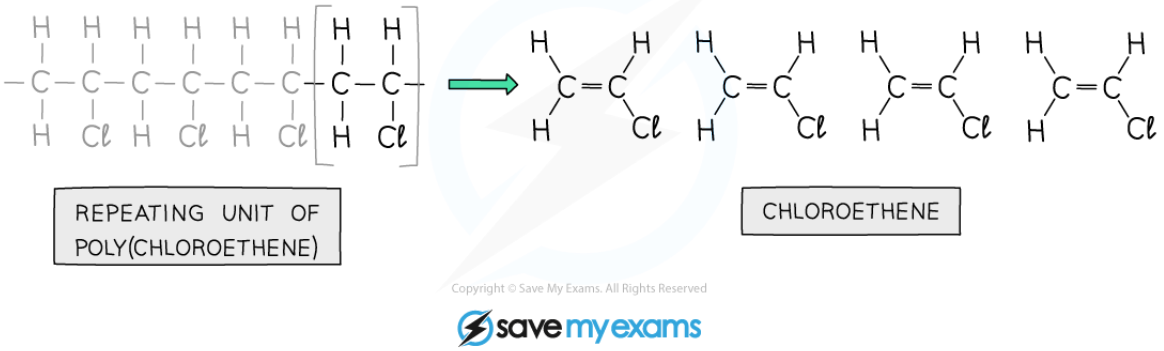


Examples of addition polymerisation: polythene and PVC

Deducing the monomer from the polymer

- Identify the repeating unit in the polymer
- Change the single bond in the repeat unit to a **double bond** in the monomer
- Remove the bond from each end of the repeat unit

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Diagram showing the monomer from the repeat unit of an addition polymer (polychloroethene)

Exam Tip

You should be able to draw the **box diagrams** representing polymers where each box represents a part of the repeating hydrocarbon chain. The **functional groups** on the monomers and the **link** formed in the polymers are the important parts and must be clearly drawn.

Your notes

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Condensation Polymers


Your notes

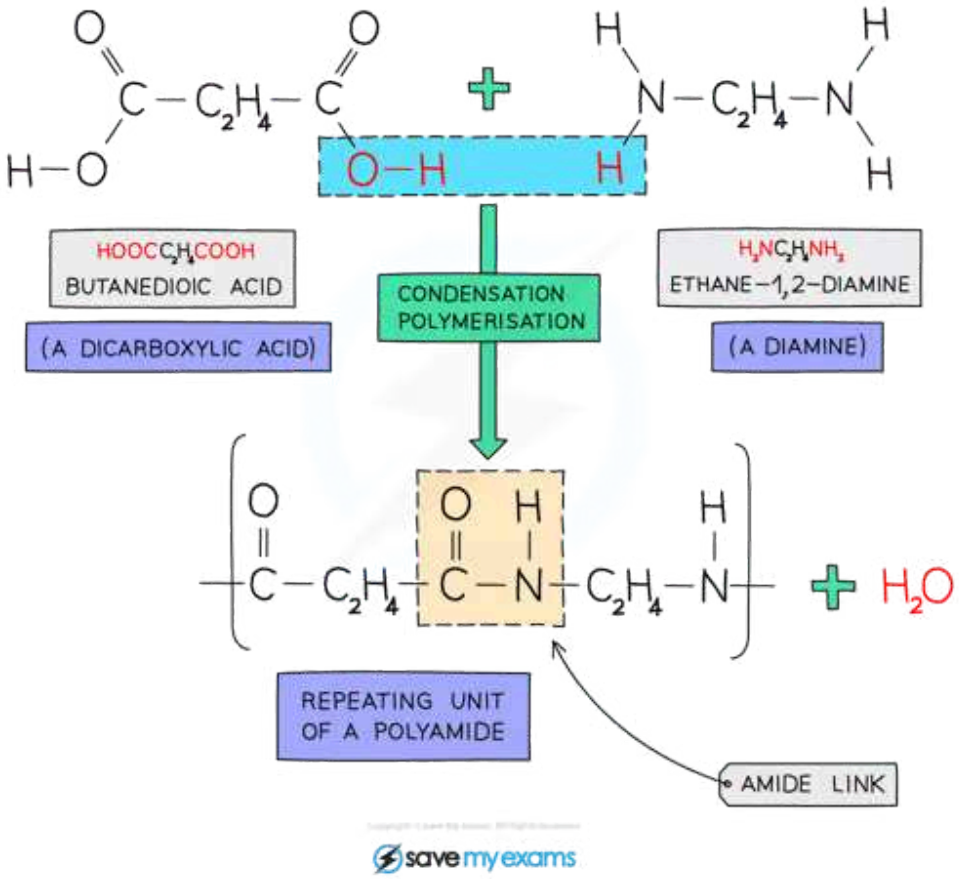
- Condensation polymers are formed when **two different monomers** are linked together with the **removal** of a small molecule, usually **water**
- This is a key difference between condensation polymers and addition polymers:
 - Addition polymerisation forms the polymer molecule **only**
 - Condensation polymerisation forms the **polymer** molecule and one **water** molecule **per linkage**
- The monomers have two **functional** groups present, one on **each end**
- The functional groups at the ends of one monomer react with the functional group on the end of the other monomer, in so doing creating long chains of **alternating monomers**, forming the polymer
- Hydrolysing (adding water) to the compound in acidic conditions usually reverses the reaction and produces the monomers by rupturing the peptide link

Forming Nylon

- Nylon is a polyamide made from **dicarboxylic** acid monomers (a carboxylic with a -COOH group at **either** end) and **diamines** (an amine with an $-NH_2$ group at **either** end)
- Each -COOH group reacts with another $-NH_2$ group on another monomer
- An amide linkage is formed with the subsequent loss of **one** water molecule per link

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Your notes



The condensation reaction in which the polyamide, nylon is produced

- The structure of nylon can be represented by drawing out the polymer using boxes to represent the carbon chains

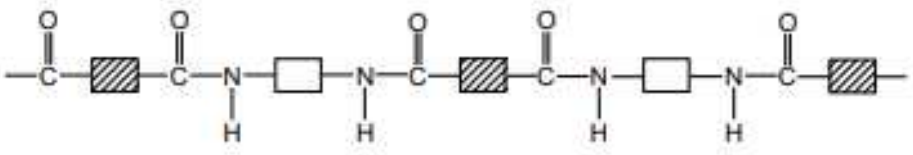


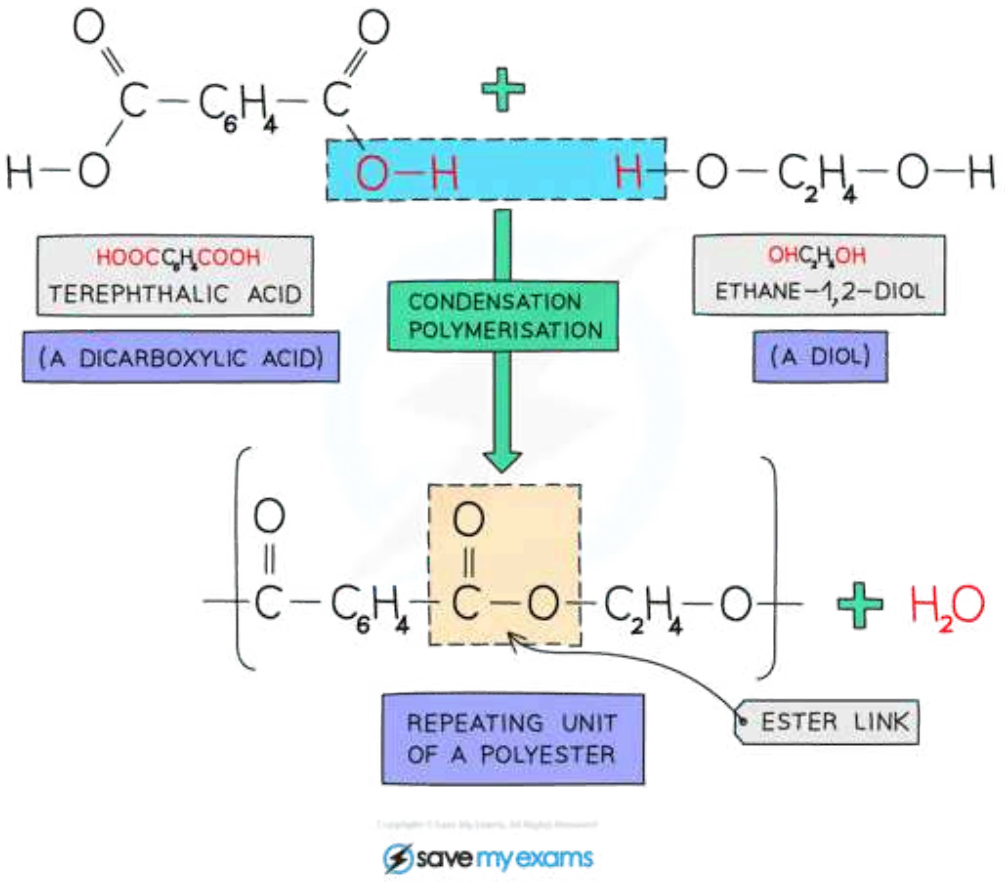
Diagram showing a section of nylon

Forming Polyesters

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Your notes

- **PET** or polyethylene terephthalate to give its full name, is a **polyester** made from **dicarboxylic** acid monomers (a carboxylic with a -COOH group at **either** end) and **diols** (alcohol with an -OH group at **either** end)
- Each -COOH group reacts with another -OH group on another monomer
- An ester linkage is formed with the subsequent loss of **one** water molecule per link
- For every ester linkage formed in condensation polymerisation, one molecule of water is formed from the combination of a **proton** (H^+) and a **hydroxyl** ion (OH^-)
- **PET** is also used in synthetic fibres as is sold under the trade name of **terylene**



The condensation reaction in which PET is produced

- The structure of PET can be represented by drawing out the polymer using boxes to represent the carbon chains
- This can be done for all polyesters

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Diagram showing a section of PET

Your notes

Exam Tip

You don't need to know the detailed chemical structure of PET, just the symbolic drawing showing the alternating blocks and the linking ester group. Be careful not to exactly repeat the linking group in nylon or PET; the link alternates by reversing the order of the atoms, rather like a mirror image.

Differences between Addition & Condensation Polymerisation

Table showing differences between addition and condensation polymerisation

Addition Polymerisation	Condensation Polymerisation
Involves **one monomer** type containing a C=C bond	Involves **two monomer** types each containing two of the same functional group **OR** One monomer type with two different functional groups
One product formed	**Two types of product** formed- the polymer, and a small molecule (usually water)
Functional group involved is a **C=C bond**	Involves **two functional groups** on each monomer

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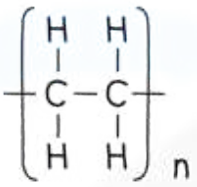
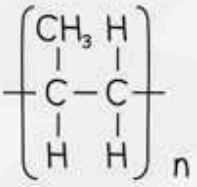
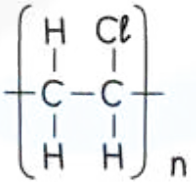
Plastics & Their Disposal

Your notes

Plastics & their Disposal

- **Synthetic** polymers are ones made in a factory, for example nylon, terylene and lycra
- Nylon is a polyamide used to produce **clothing**, **fabrics**, **nets** and **ropes**
- PET, also known as Terylene, is a polyester made from monomers which are joined together by **ester links**
- PET is used extensively in the **textile** industry and is often mixed with cotton to produce **clothing**

Table showing Uses of Plastics

POLYMER	REPEAT UNIT	USES
POLY(ETHENE) I.e. POLYTHENE		PLASTIC BAGS (LOW DENSITY POLYTHENE) PLASTIC BOTTLES (HIGH DENSITY POLYTHENE)
POLY(PROPENE)		FOOD PACKAGING ROPES CARPETS
POLY(CHLOROETHENE) I.e. PVC		PLASTIC SHEETS ARTIFICIAL LEATHER DRAINPIPES AND GUTTERS INSULATION ON WIRES

Non-biodegradable plastics

- These are plastics which do not degrade over time or take a very long time to degrade, and cause significant pollution problems
- In particular plastic waste has been spilling over into the **seas** and **oceans** and is causing huge disruptions to marine life

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- In landfills waste polymers take up valuable space as they are non-biodegradable so microorganisms cannot break them down. This causes the landfill sites to quickly fill up
- Polymers release a lot of heat energy when **incinerated** and produce **carbon dioxide** which is a greenhouse gas that contributes to climate change
- If incinerated by incomplete combustion, **carbon monoxide** will be produced which is a toxic gas that reduces the capacity of the blood to carry oxygen
- Polymers can be recycled but different polymers must be separated from each other which is a difficult and expensive process

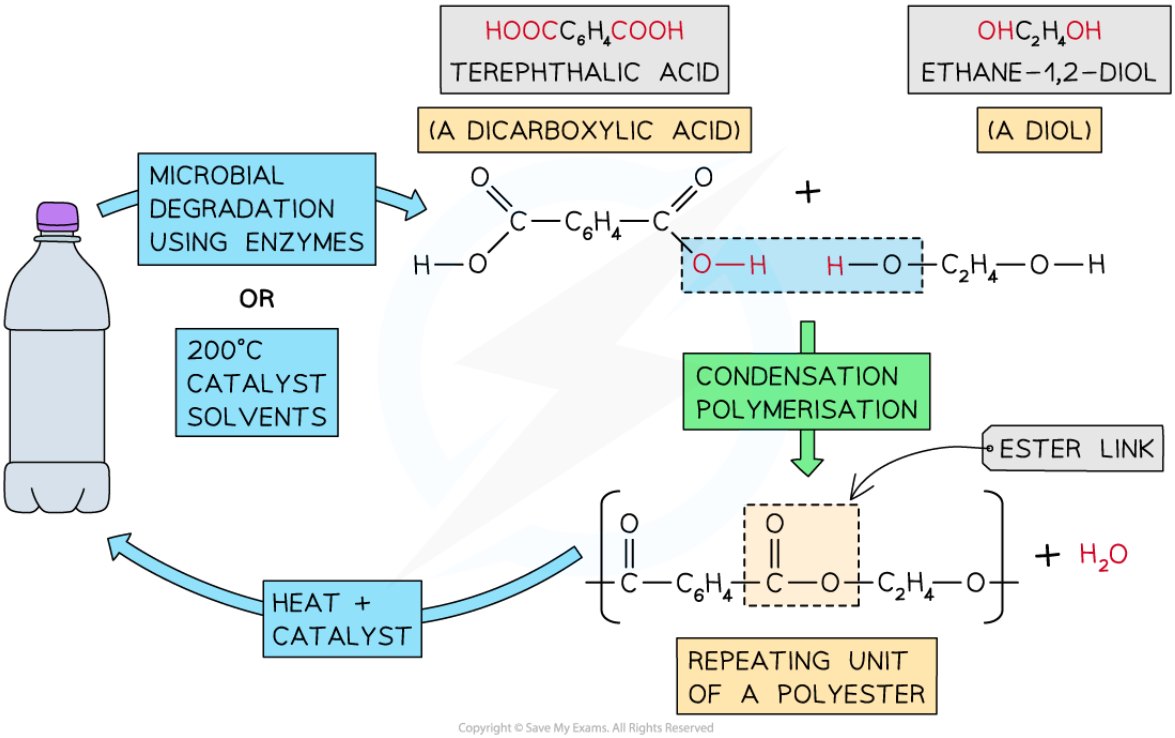

Your notes

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PET Re-polymerisation

Your notes

- **PET** stands for polyethylene terephthalate, a common polymer used to make things like plastic bottles
- It is a **condensation** polymer consisting of repeating ester units, so it is a type of polyester
- One of the problems with recycling polymers is that the conditions needed to break them down, which are usually high temperatures and pressures, can degrade the monomers making them unusable for re-polymerisation
- PET is relatively easy to convert back into the monomers
- It can be depolymerised either using enzymes or by chemical methods
- Enzymes present in microbes breakdown the PET into the original monomers
- The same can be achieved using solvents a catalyst and mild heating



The breakdown of PET into its two monomers takes place using enzymes or chemical catalysts and mild conditions

- The monomers are recovered and polymerised into new PET
- This saves on resources and energy, reducing the carbon footprint of the production process

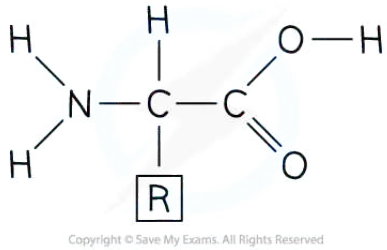
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Proteins

Your notes

Proteins

- Proteins are **condensation polymers** which are formed from amino acid monomers joined together by **amide** links (in proteins also known as a peptide link) similar to the structure in nylon
- The units in proteins are different however, consisting of amino acids
- Amino acids are small molecules containing **NH_2** and **COOH** functional groups



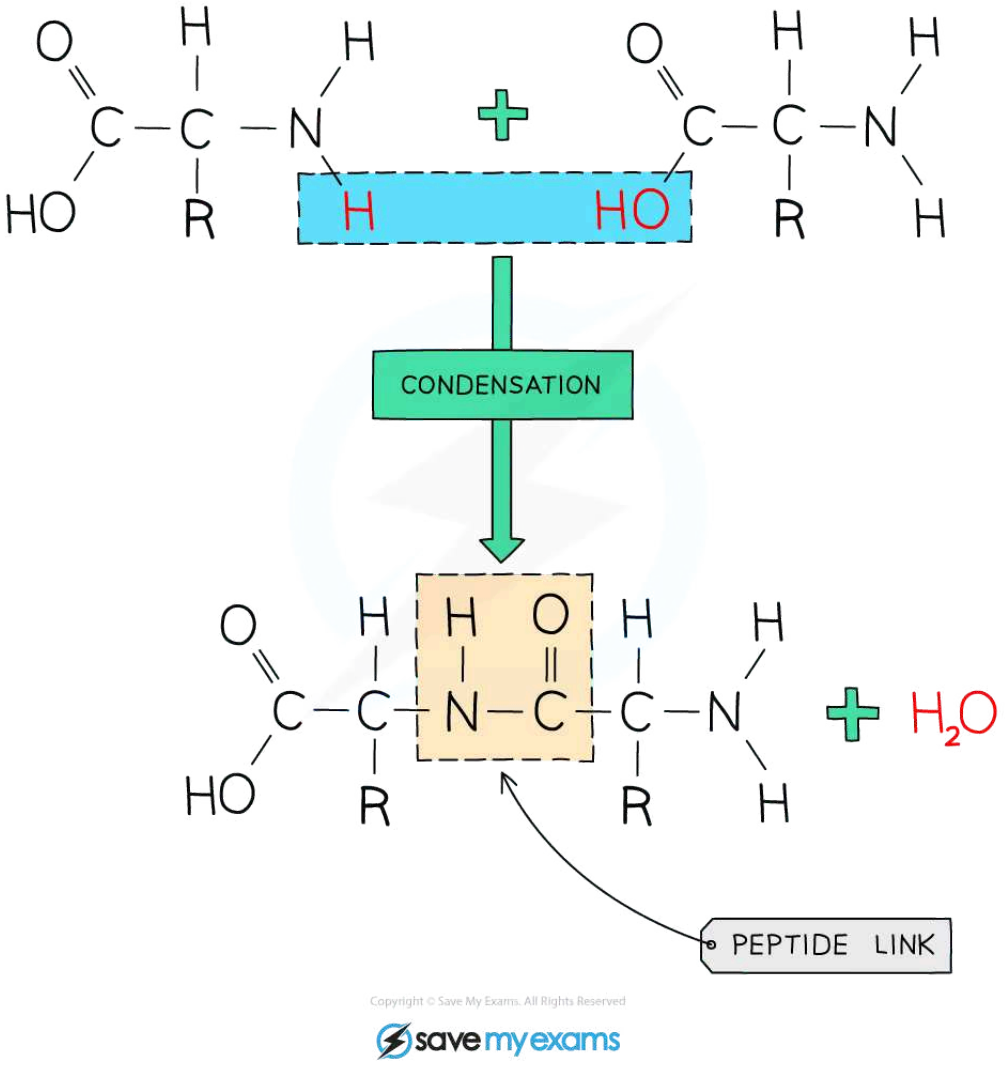
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General structure of an amino acid

- There are **twenty** common amino acids, each differing by their side chain, represented by R
- Proteins can contain between 60 and 600 of these amino acids in different orders
- These are the monomers which polymerise to form the protein

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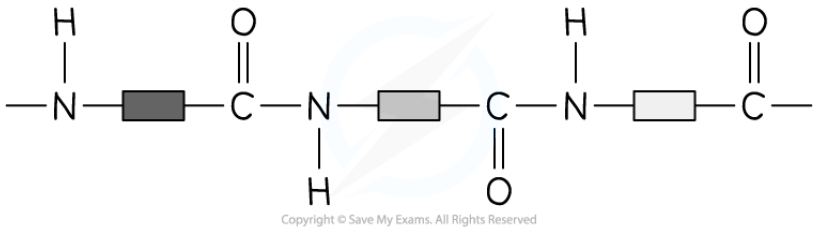


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Diagram showing condensation polymerisation to produce a protein

- The structure of proteins can be represented using the following diagram whereby the boxes represent the carbon chains

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Diagram showing a section of protein

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Cambridge O Level Chemistry

Your notes

12.1 Experimental Design

Contents

- Apparatus for Measurements
- Solutions

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Apparatus for Measurements


Your notes

Time, Temperature, Mass & Volume

Time

- Time can be measured using a stopwatch or stopclock which are usually accurate to one or two decimal places
- The units of time normally used are seconds or minutes although other units may be used for extremely slow reactions (e.g. rusting)
- 1 minute = 60 seconds

Temperature

- Temperature is measured with a thermometer or digital probe
- Laboratory thermometers usually have a precision of a half or one degree
- Digital temperature probes are available which are more precise than traditional thermometers and can often read to 0.1 °C
- Traditional thermometers rely upon the uniform expansion and contraction of a liquid substance with temperature; digital temperature probes can be just as, if not, more accurate than traditional thermometers
- The units of temperature are **degrees Celsius (°C)**

Mass

- Mass is measured using a digital balance which normally gives readings to two decimal places
- Balances must be tared (set to zero) before use
- The standard unit of mass in kilograms (kg) but in chemistry grams (g) are most often used
- 1 kilogram = 1000 grams

Volume-liquids

- The volume of a liquid can be determined using several types of apparatus, depending on the level of accuracy needed
- For approximate volumes where high accuracy is not an important factor, measuring (or graduated) cylinders are used
- These are graduated (have a scale so can be used to measure) and are available typically in a range of sizes from 10 cm^3 to 1 litre (1 dm^3)
- Volumetric pipettes are the most accurate way of measuring a **fixed** volume of liquid, usually 10 cm^3 or 25 cm^3
 - They have a scratch mark on the neck which is matched to the bottom of the meniscus to make the measurement
- Burettes are the most accurate way of measuring a **variable** volume of liquid between 0 cm^3 and 50 cm^3 (e.g. in a titration)

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- The tricky thing with burettes is to remember to read the scale from top to bottom as 0.00 cm^3 is at the top of the column
- Whichever apparatus you use, you may see markings in ml(millilitre) which is the same as a cm^3

Your notes

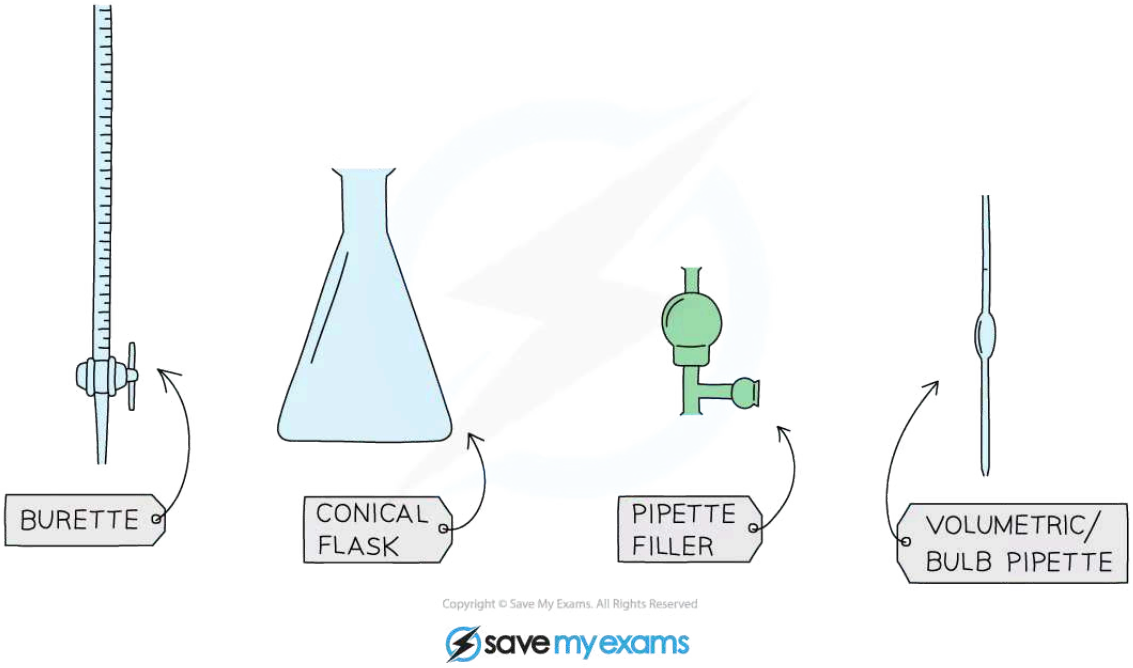


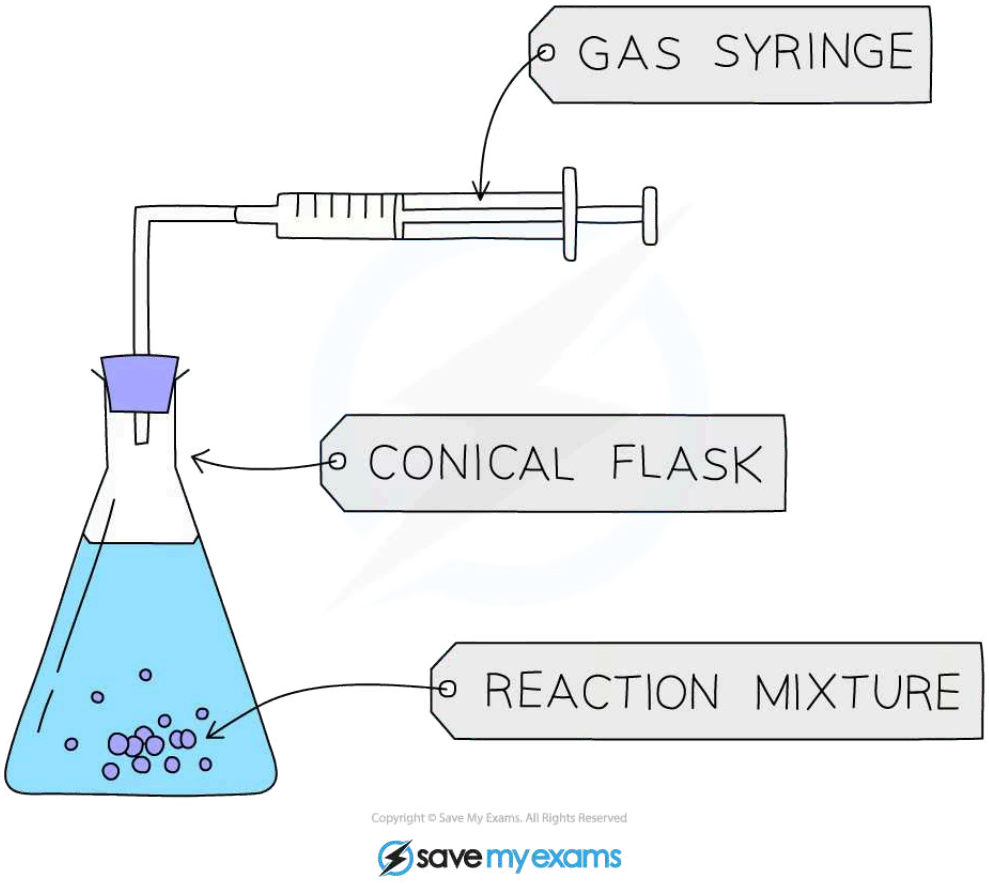
Diagram of a burette with conical flask and pipette with filler

Volume-gases

- The volume of a gas sometimes needs to be measured and is done by collecting it in a graduated measuring apparatus
- A **gas syringe** is usually the apparatus used
- A graduated cylinder inverted in water may also be used, provided the gas isn't water-soluble
- If the gas happens to be heavier than air and is coloured, the cylinder can be used upright

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Your notes



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Diagram of the set-up for an experiment involving gas collection

Exam Tip

Be careful when recording time not to mix up seconds and minutes in the same table. If a table heading shows Time/mins and you record a stop watch display of 1.30, meaning 1 minute and 30 seconds, that is wrong as it should be 1.5 mins. To avoid any confusion, if the time intervals are less than a minute, its best to change the recorded units to seconds. That same stopwatch display would therefore be recorded as 90 seconds.

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Advantages & Disadvantages of Methods & Apparatus


Your notes

- In the lab we often have choices of different apparatus to do the same job
- Evaluating what is the best one to use is part of good experimental planning and design
- This means appreciating some of the advantages and disadvantages of laboratory apparatus

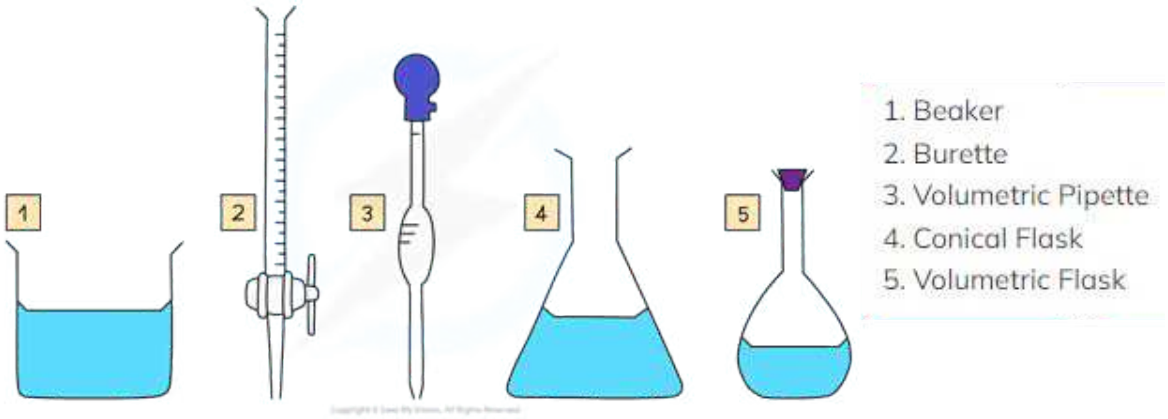
Table showing examples of Advantages and Disadvantages of Lab Apparatus

Apparatus choice	Advantage	Disadvantage
Temperature probe versus thermometer	◦ More precise readings ◦ Can easily take multiple repeat readings ◦ Can use automatic sampling over longer periods of time	◦ Probe can be corroded by some reagents ◦ Probes more expensive to replace
Volumetric pipette versus a measuring cylinder	◦ Pipette measures very accurately	◦ Harder to use and only measures one fixed volume
Gas syringe versus inverted cylinder for collecting gases	◦ Gas syringes are easy to set up and keep the gas dry	◦ The pistons can stick ◦ Limited volume can be collected ◦ Delicate and expensive
Mircroscale versus normal scale quantities	◦ Less wasteful of reagents ◦ Saves energy ◦ Safer	◦ Hard to see what's happening ◦ Lose a lot of material in separating and purifying products

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Your notes



Five pieces of apparatus that can be used to measure the volume of a liquid. They all have their pros and cons

Planning your method

- Good experimental design includes the answers to questions like
 - Have I chosen a suitable apparatus for what I need to measure?
 - Is it going to give me results in an appropriate time frame?
 - Is it going to give me enough results to process, analyse and make conclusions?
 - Does it allow for repetitions to check how **reliable** my results are?
 - Does my plan give a suitable **range** of results?
 - How can I be sure my results are **accurate**?
 - Have I chosen an appropriate scale of quantities without being wasteful or unsafe?
- You may be asked about experimental methods in exam questions and your experience and knowledge of practical techniques in chemistry should help you to spot mistakes and suggest improvements

Exam Tip

Make sure you know the names of common laboratory apparatus

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Solutions

Your notes

Solutions

- You need to know all the following terms used when describing solutions:

Terminology About Solutions Table

Term	Meaning	Example
Solvent	The liquid in which a solute dissolves	The water in seawater
Solute	The substance which dissolves in a liquid to form a solution	The salt in seawater
Solution	The mixture formed when a solute is dissolved in a solvent	Seawater
Saturated solution	A solution with the maximum concentration of solute dissolved in the solvent	Seawater in the dead sea
Soluble	Describes a substance that will dissolve	Salt is soluble in water
Insoluble	Describes a substance that won't dissolve	Sand is insoluble in water
Filtrate	The liquid or solution that has passed through a filter	Fresh coffee in a cup
Residue	A substance that remains after evaporation, distillation, filtration or any other similar process	Coffee grounds in filter paper

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Cambridge O Level Chemistry

Your notes

12.2 Acid-Base Titrations

Contents

* Acid-Base Titrations

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Acid-Base Titrations


Your notes

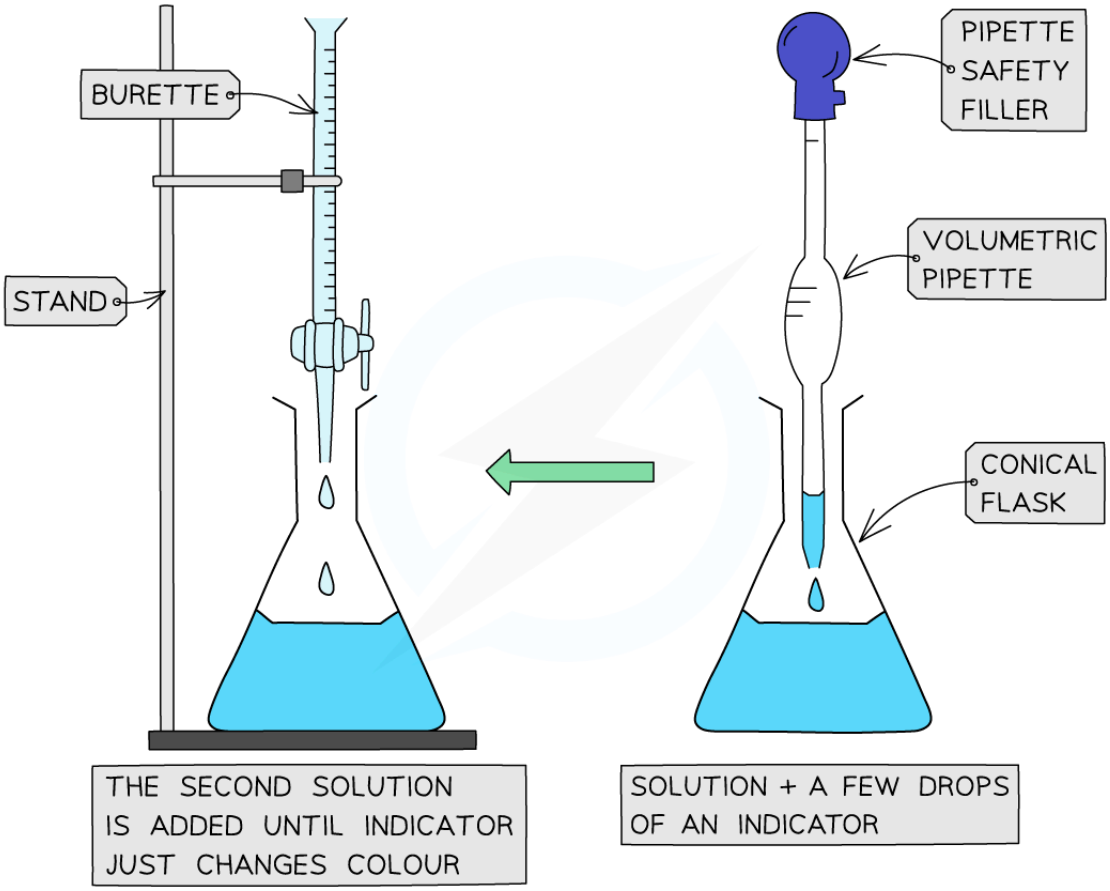
Acid-Base Titrations

- Titrations are a method of analysing the **concentration** of solutions
- They can determine exactly how much alkali is needed to neutralise a quantity of acid – and vice versa
- You may be asked to calculate the **moles** present in a given amount, the **concentration** or **volume** required to **neutralise** an acid or a base
- Titrations can also be used to prepare **salts**

Describing how to carry out a titration

- **The typical materials needed are:**
 - 25 cm^3 volumetric pipette
 - Pipette filler
 - 50 cm^3 burette
 - 250 cm^3 conical flask
 - Small funnel
 - 0.1 mol/dm^3 sodium hydroxide solution
 - Sulfuric acid – concentration unknown
 - A suitable indicator
 - Clamp stand, clamp & white tile

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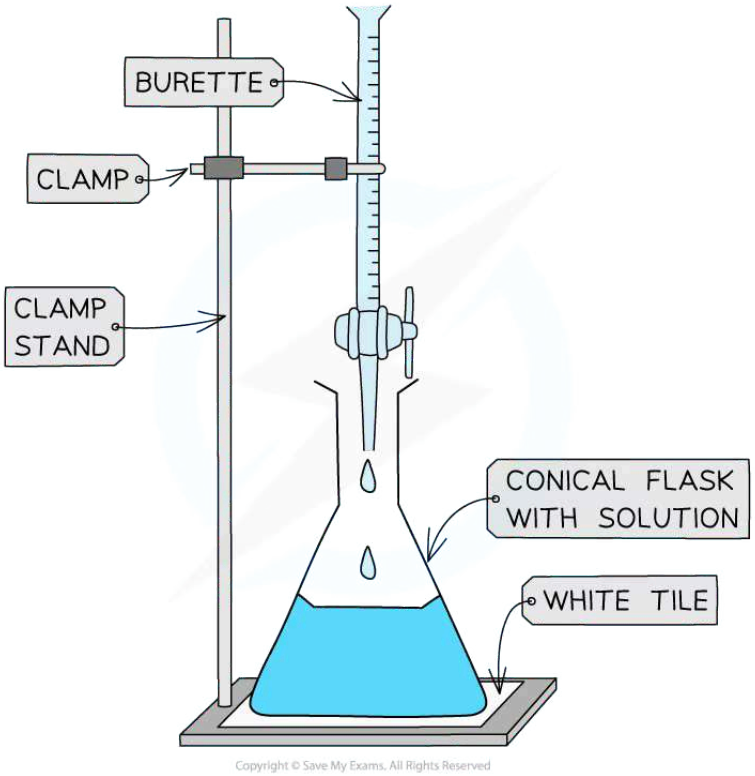


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Your notes



The steps in performing a titration

Describing the method:

1. Use the pipette and pipette filler and place exactly 25 cm^3 sodium hydroxide solution into the conical flask
2. Place the conical flask on a white tile so the tip of the burette is inside the flask
3. Add a few drops of a suitable indicator to the solution in the conical flask
4. Perform a rough titration by taking the burette reading and running in the solution in 1 – 3 cm^3 portions, while swirling the flask vigorously
5. Quickly close the tap when the end-point is reached (sharp colour change) and record the volume, placing your eye level with the meniscus
6. Now repeat the titration with a fresh batch of sodium hydroxide
7. As the rough end-point volume is approached, add the solution from the burette one drop at a time until the indicator just changes colour
8. Record the volume to the nearest 0.05 cm^3
9. Repeat until you achieve two concordant results (two results that are within 0.1 cm^3 of each other) to increase accuracy

Indicators

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- **Indicators** are used to show the endpoint in a titration
- Wide range indicators such as litmus are not suitable for titration as they do not give a sharp enough colour change at the endpoint
- However, methyl orange and phenolphthalein are very suitable
- Some of the most common indicators with their corresponding colours are shown below:


Your notes

Common Acid-Base Indicators

Indicator	Colour in acid	Colour in alkaline	Colour in neutral
Litmus solution	Red	Blue	Purple
Litmus paper : red	Stays red	Turns blue	No change
Litmus paper : blue	Turns red	Stays blue	No change
Methyl orange	Red	Yellow	Orange
Phenolphthalein	Colourless	Pink	Colourless
Thymolphthalein	Colourless	Blue	Colourless

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Cambridge O Level Chemistry

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12.3 Chromatography

Contents

- Paper Chromatography
- Locating Agents & Rf Values

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Paper Chromatography


Your notes

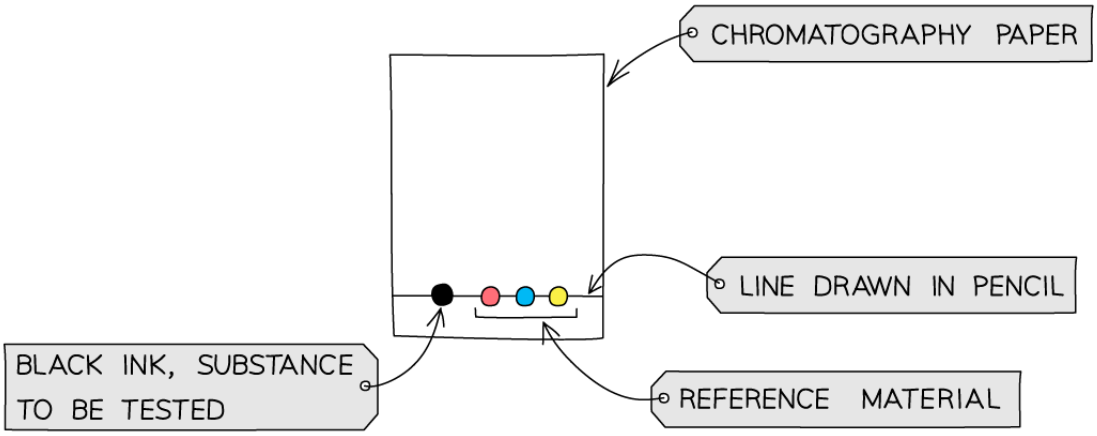
Paper Chromatography

- This technique is used to separate substances that have **different solubilities** in a given solvent (e.g. different coloured inks that have been mixed to make black ink)
- A **pencil line** is drawn on chromatography paper and spots of the sample are placed on it. Pencil is used for this as ink would run into the chromatogram along with the samples
- The paper is then lowered into the solvent container, making sure that the pencil line sits **above** the level of the solvent so the samples don´t wash into the solvent container
- The solvent travels up the paper by capillary action, taking some of the coloured substances with it
- Different substances have different solubilities so will travel at different rates, causing the substances to spread apart. Those substances with higher solubility will travel further than the others
- This will show the different components of the ink / dye

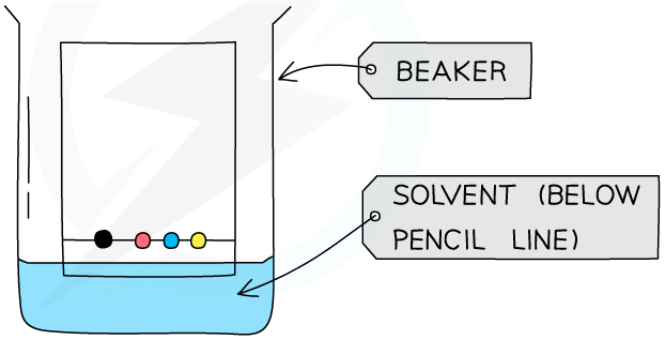
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Your notes

1 SET UP CHROMATOGRAPHY PAPER AS SHOWN

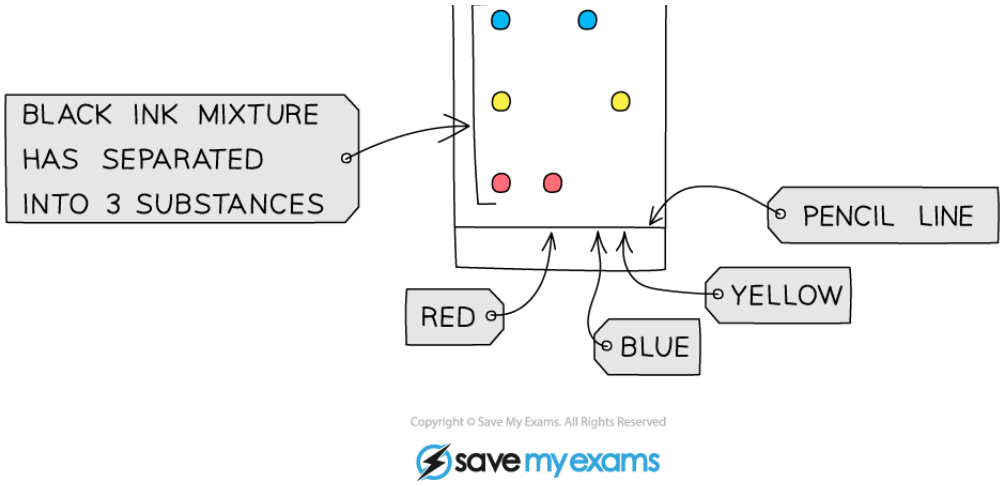


2 LOWER PAPER INTO A BEAKER WITH APPROPRIATE SOLVENT. WAIT FOR SOLVENT TO TRAVEL UP THE PAPER.



3 ANALYSE CHROMATAGRAM

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Analysis of the composition of ink using paper chromatography

Interpret Simple Chromatograms

- If two or more substances are the same, they will produce identical chromatograms
- It is common practice to include a known compound as a reference spot
 - This can help match up to an unknown spot or set of spots in order to identify it
- If the substance is a mixture, it will separate on the paper to show all the different components as separate spots
- An impure substance will show up with more than one spot, a pure substance should only show up with one spot

Exam Tip

Paper chromatography is the name given to the **overall** separation technique while a **chromatogram** is the name given to the **visual output** of a chromatography run. This is the piece of chromatography paper with the visibly separated components after the run has finished.

Your notes

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Locating Agents & Rf Values


Your notes

Locating Agents

- For chromatography to be useful the chemist needs to be able to see the components move up the paper, which is not the case for colourless substances such as amino acids or sugars
- Locating agents are substances which react with the sample and produce a coloured product which is then visible
- The chromatogram is treated with the agent **after** the chromatography run has been carried out, making the sample runs visible to the naked eye

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Retention Factor (Rf) Values

Your notes

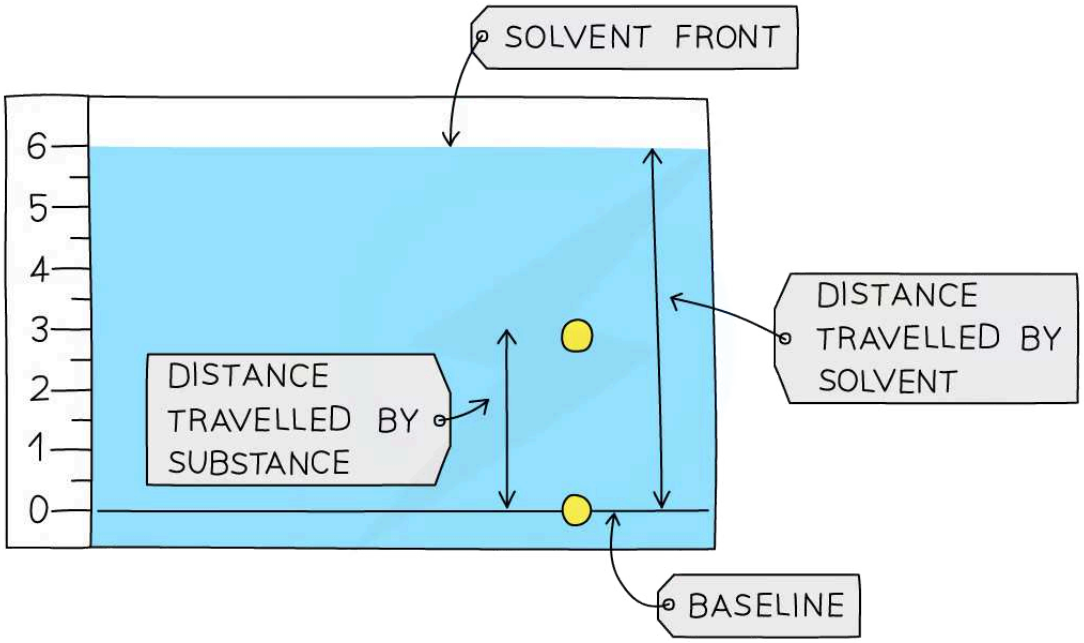
- These values are used to identify the components of mixtures
- The R_f value of a particular compound is always the same
- Calculating the R_f value allows chemists to identify unknown substances because it can be compared with R_f values of known substances under the same conditions

Calculation

- The formula to calculate the retention factor is:

$$\text{Retention factor} = \frac{\text{Distance travelled by substance}}{\text{Distance travelled by solvent}}$$

- The R_f value is a ratio and therefore has no units



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Using R_f values to identify components of a mixture

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12.4 Separation & Purification

Contents

Separation & Purification Techniques

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Separation & Purification Techniques


Your notes

Filtration & Crystallisation

- The choice of the method of separation depends on the nature of the substances being separated
- All methods rely on there being a **difference** of some sort, usually in a physical property such as boiling point between the substances being separated

Separating a mixture of solids

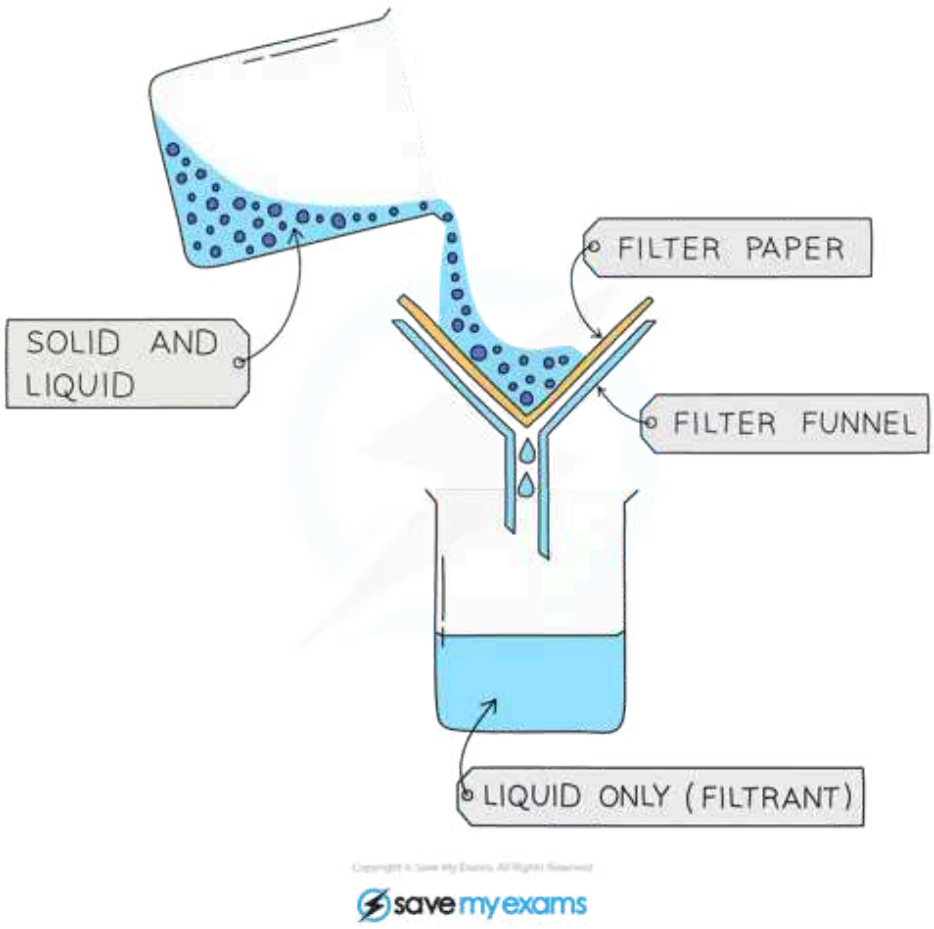
- Differences in solubility can be used to separate solids
- For a difference in solubility, a suitable **solvent** must be chosen to ensure the desired substance only dissolves in it and not other substances or impurities, e.g. to separate a mixture of sand and salt, water is a suitable solvent to dissolve the salt, but not the sand

Filtration

- Used to separate an **undissolved solid** from a mixture of the solid and a liquid / solution (e.g. sand from a mixture of sand and water). Centrifugation can also be used for this mixture
- Filter paper is placed in a filter funnel above another beaker
- The mixture of insoluble solid and liquid is poured into the filter funnel
- Filter paper will only allow small liquid particles to pass through in the filtrate
- Solid particles are too large to pass through the filter paper so will stay behind as a residue

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Filtration of a mixture of sand and water

Crystallisation

- Used to separate a **dissolved solid** from a solution, when the solid is more soluble in hot solvent than in cold (e.g. copper sulphate from a solution of copper (II) sulphate in water)
- The solution is heated, allowing the solvent to evaporate and leaving a saturated solution behind
- You can test if the solution is saturated by dipping a clean, dry, cold glass rod into the solution
 - If the solution is saturated, crystals will form on the glass rod when it is removed and allowed to cool
- The saturated solution is allowed to cool slowly and solids will come out of the solution as the solubility decreases, and crystals will grow
- Crystals are collected by filtering the solution
- They are then washed with distilled water to remove any impurities and then allowed to dry

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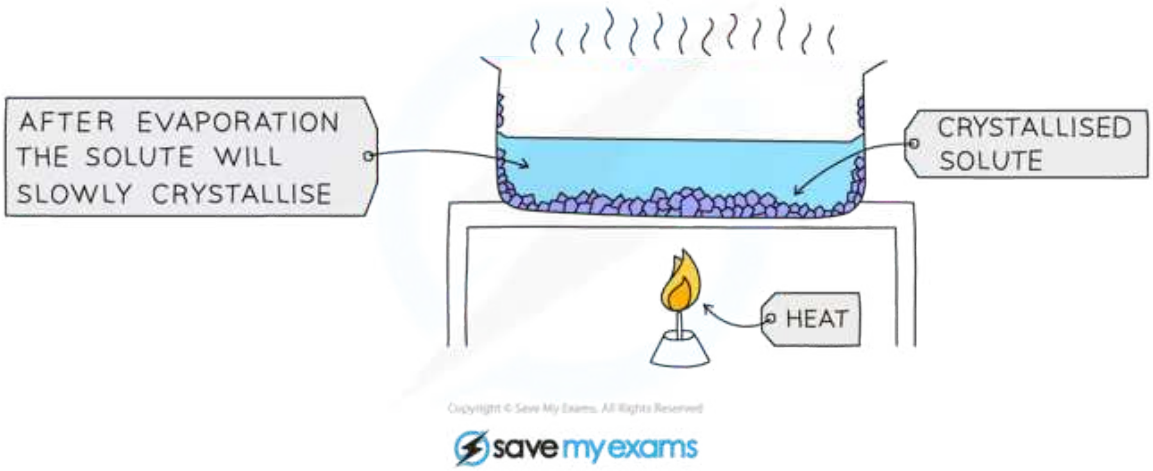


Diagram showing the process of crystallisation

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Distillation: Simple & Fractional

Your notes

Simple Distillation

- Used to separate a liquid and **soluble solid** from a solution (e.g. water from a solution of saltwater) or a pure liquid from a mixture of liquids
- The solution is heated and pure water evaporates producing a vapour which rises through the neck of the round-bottomed flask
- The vapour passes through the condenser, where it cools and condenses, turning into pure water which is collected in a beaker
- After all the water is evaporated from the solution, only the solid solute will be left behind

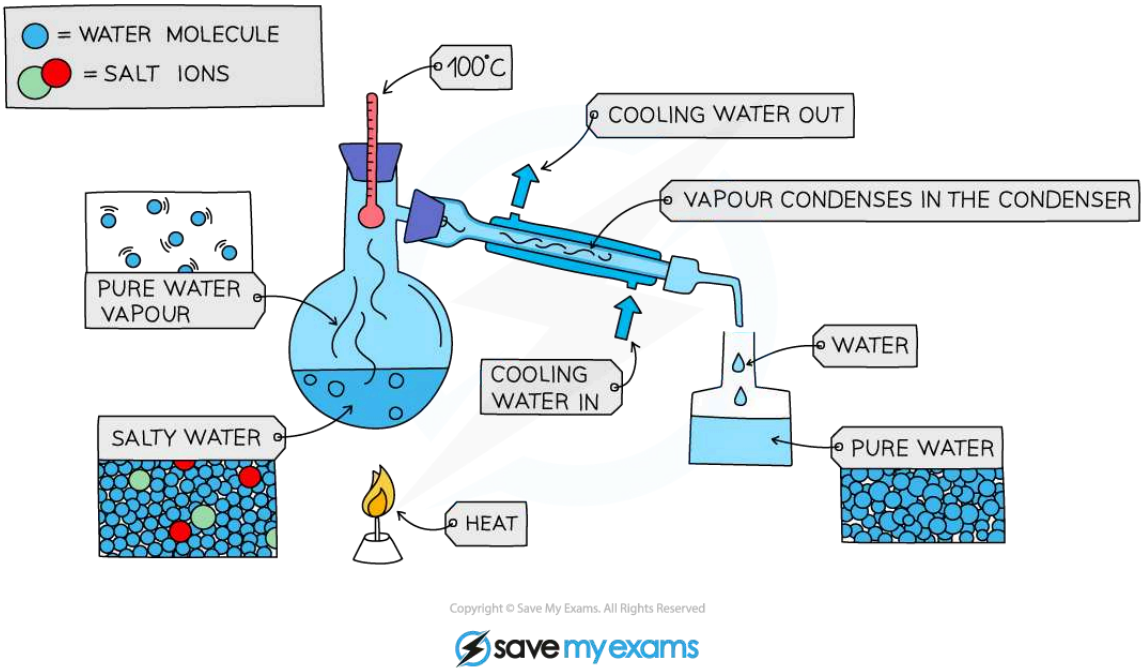


Diagram showing the distillation of a mixture of salt and water

- Simple distillation can be used to separate the products of fermentation, such as alcohol and water
- However, more effective separation is to use fractional distillation where the liquids are closer to boiling point and a higher degree of purity is required

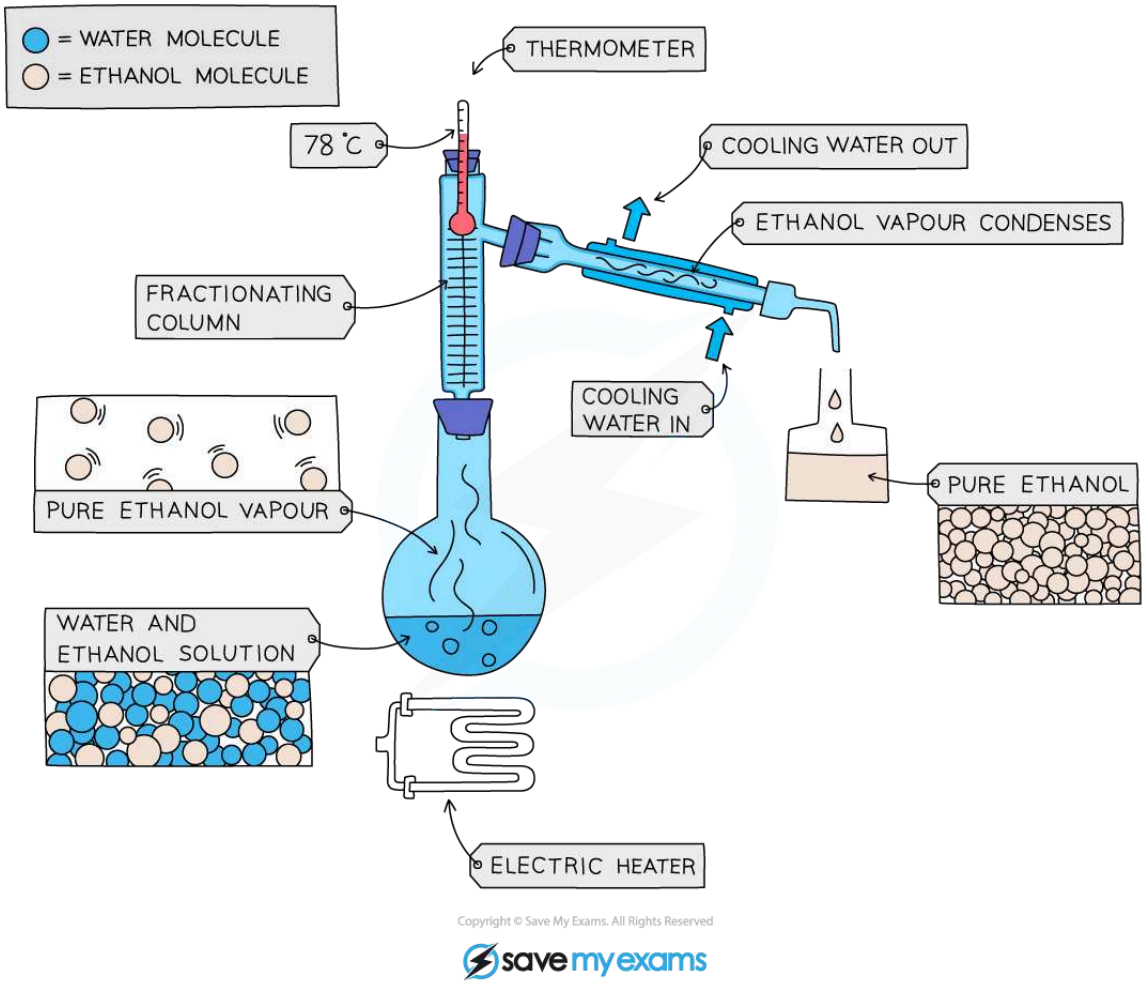
Fractional distillation

- Used to separate two or more liquids that are miscible with one another (e.g. ethanol and water from a mixture of the two)
- The solution is heated to the temperature of the substance with the lowest boiling point

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Your notes

- This substance will rise and evaporate first, and vapours will pass through a condenser, where they cool and condense, turning into a liquid that will be collected in a beaker
- All of the substance is evaporated and collected, leaving behind the other components(s) of the mixture
- For water and ethanol: ethanol has a boiling point of 78 °C and water of 100 °C. The mixture is heated until it reaches 78 °C, at which point the ethanol boils and distills out of the mixture and condenses into the beaker
- When the temperature starts to increase to 100 °C heating should be stopped. Water and ethanol are now separated



Fractional distillation of a mixture of ethanol and water

- An electric heater is safer to use when there are flammable liquids present

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- The separation of the components in petroleum is achieved by fractional distillation on an industrial scale
- Fractional distillation of crude oil is not carried out in school laboratories due to the toxic nature of some of the components of the crude oil, but it can sometimes be simulated using a synthetic crude oil made specially for the demonstration

Your notes

Worked example

A student is given a mixture of calcium sulfate, magnesium chloride and water. The table below shows some information about calcium sulfate and magnesium chloride.

substance	solubility in water	state at room temperature
calcium sulfate	insoluble	solid
magnesium chloride	soluble	solid

How does the student obtain magnesium chloride crystals from the mixture?

A crystallisation followed by distillation
B crystallisation followed by filtration
C distillation followed by crystallisation
D filtration followed by crystallisation

Answer:

The correct answer is **D** because:

- The difference in solubility in water means the first step is to make a solution
- The magnesium chloride will dissolve, but the solid calcium sulfate will be left behind
- The mixture is filtered to remove the calcium sulfate and then evaporated and crystallised to obtain magnesium chloride crystals

Exam Tip

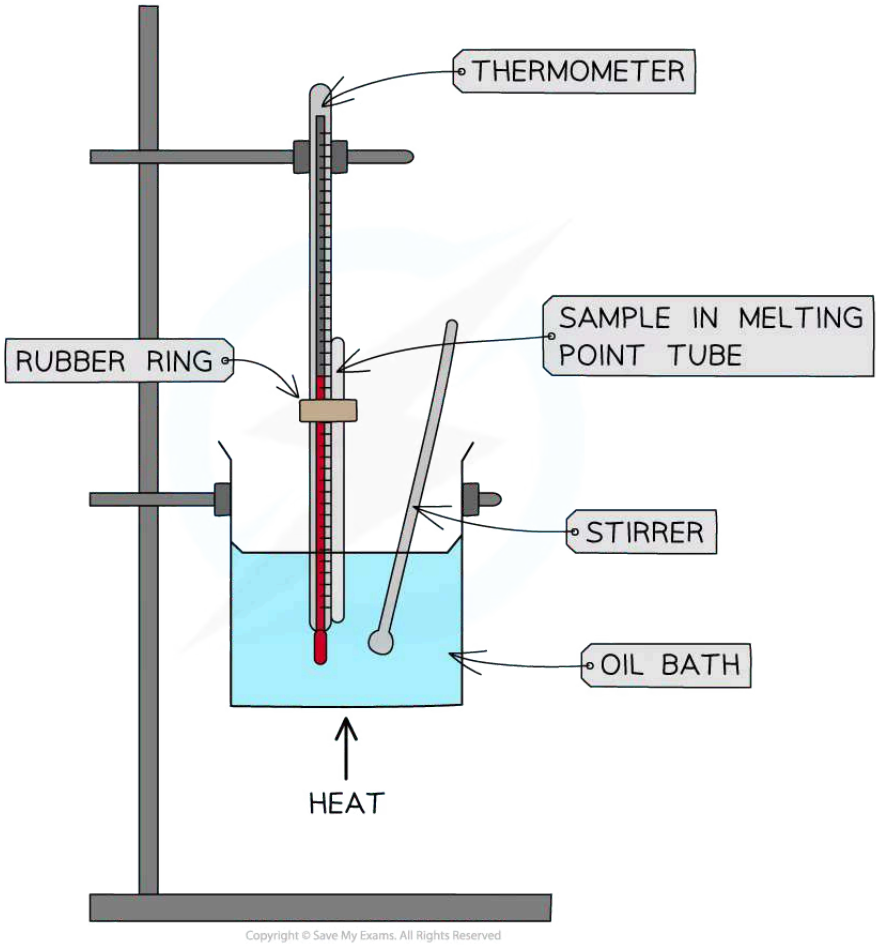
You may be asked how to separate a mixture of gases. One method involves cooling the gaseous mixture sufficiently to liquefy all of the gases. The liquified mixture is then separated by fractional distillation. They can also be separated by diffusion, where the boiling points are very close or it is impractical or expensive to use fractional distillation.

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Assessing Purity

Your notes

- Pure substances melt and boil at specific and sharp temperatures. Eg water has a boiling point of 100°C and a melting point of 0°C
- Mixtures have a range of melting and boiling points as they consist of different substances that melt or boil at different temperatures
- Melting and boiling point data can therefore be used to distinguish pure substances from mixtures
- An unknown pure substance can be identified by experimentally determining its melting point and boiling point and comparing them to literature values / data tables
 - Boiling points are commonly determined by distillation
- Melting point analysis is routinely used to assess the purity of drugs for example
- This is done using a melting point apparatus which allows you to slowly heat up a small amount of the sample, making it easier to observe the exact melting point



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Melting point test using an oil bath


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- This is then compared to data tables
- The closer the measured value is to the actual melting or boiling point then the purer the sample is
- If the sample contains impurities:
 - The boiling point may appear higher than the sample's actual boiling point
 - The melting point may appear lower than the sample's actual boiling point

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Cambridge O Level Chemistry

Your notes

12.5 Identification of Ions & Gases

Contents

- Identification of Anions
- Identification of Cations
- Identification of Gases

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Identification of Anions

Your notes

Identification of Anions

Anion	Test	Result
Carbonate (CO_3^{2-})	Add dilute acid and test the gas released	Effervescence, gas produced is CO_2 which turns limewater milky
Chloride (Cl^-)	Acidify with dilute nitric acid and add aqueous silver nitrate	White precipitate formed
Bromide (Br^-)	Acidify with dilute nitric acid and add aqueous silver nitrate	Cream precipitate formed
Iodide (I^-)	Acidify with dilute nitric acid and add aqueous silver nitrate	Yellow precipitate formed
Nitrate (NO_3)	Add aqueous NaOH and aluminium foil, warm gently and test the gas released	Gas given off is ammonia, has a pungent smell and turns moist red litmus paper blue
Sulfate (SO_4^{2-})	Acidify with dilute nitric acid and add aqueous barium nitrate	White precipitate formed
Sulfite (SO_3^{2-})	Add dilute acid, warm gently and test the gas released	Gas decolourises purple acidified aqueous potassium manganate(VII) solution

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- The acidification with aqueous silver nitrate and aqueous barium nitrate / chloride is done to remove ions that might give a false positive result
 - The most common ion that is removed is the carbonate ion
 - The choice of acid has to be considered so that it does not influence the results

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- For example, aqueous silver nitrate solution should not be acidified with hydrochloric acid as this will form a white precipitate due to the chloride ion in the acid
- For example, aqueous barium nitrate / chloride solution should not be acidified with sulfuric acid as this will form a white precipitate due to the sulfate ion in the acid



Your notes

Exam Tip

When it comes to qualitative inorganic analysis, always remember that there will be a test for the metal **cation** part of the molecule and another test for the **anion** part.

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Identification of Cations


Your notes

Identification of Cations

Identification of Cations

- Metal cations in aqueous solution can be identified by the **colour** of the precipitate they form on addition of sodium hydroxide and ammonia
- If only a small amount of NaOH is used then normally the resulting **metal hydroxide** precipitates out of solution
- In **excess** NaOH some of the precipitates may re-dissolve
- A few drops of NaOH is added at first and any colour changes or precipitates formed are noted
- Then the NaOH is added in excess and the reaction is observed again
- The steps are then repeated for the test using ammonia solution

Analysing results

- The table below contains the results for each of the cations included in the syllabus
- If a precipitate is formed from either NaOH or aqueous ammonia then it means that the hydroxide is insoluble in water
- Zinc chloride, for example, reacts as such:

$$ZnCl_2(aq) + 2NaOH(aq) \rightarrow Zn(OH)_2(s) + 2NaCl(aq)$$

- Ca^{2+} ions can be distinguished from Zn^{2+} and Al^{3+} as calcium hydroxide precipitate **does not dissolve** in excess NaOH but both zinc hydroxide and aluminium hydroxide do
- Zn^{2+} ions can be distinguished from Al^{3+} ions as $Zn(OH)_2$ **dissolves** in excess aqueous ammonia but $Al(OH)_3$ **does not**
- Most transition metals produce hydroxides with distinctive colours

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Your notes

METAL CATION	EFFECT OF ADDING NaOH	EFFECT OF ADDING AMMONIA SOLUTION
ALUMINIUM (Al^{3+})	WHITE PRECIPITATE, DISSOLVES IN EXCESS NaOH TO FORM A COLOURLESS SOLUTION	WHITE PRECIPITATE, INSOLUBLE IN EXCESS AMMONIA, WHITE PRECIPITATE REMAINS
AMMONIUM (NH^{4+})	AMMONIA PRODUCED IF WARMED	–
CALCIUM (Ca^{2+})	WHITE PRECIPITATE, INSOLUBLE SO REMAINS IN EXCESS NaOH	VERY FAINTLY VISIBLE WHITE PRECIPITATE
CHROMIUM (III) (Cr^{3+})	GREEN PRECIPITATE WHICH FORMS A GREEN SOLUTION IN EXCESS	GREY–GREEN PRECIPITATE, INSOLUBLE IN EXCESS
COPPER (II) (Cu^{2})	LIGHT BLUE PRECIPITATE, INSOLUBLE IN EXCESS	LIGHT BLUE PRECIPITATE, SOLUBLE IN EXCESS TO FORM DARK BLUE COLOUR
IRON (II) (Fe^{2+})	GREEN PRECIPITATE, INSOLUBLE IN EXCESS	GREEN PRECIPITATE, INSOLUBLE IN EXCESS
IRON (III) (Fe^{3+})	RED–BROWN PRECIPITATE, INSOLUBLE IN EXCESS	RED–BROWN PRECIPITATE, INSOLUBLE IN EXCESS
ZINC (Zn^{2+})	WHITE PRECIPITATE, DISSOLVES IN EXCESS TO FORM COLOURLESS SOLUTION	WHITE PRECIPITATE, DISSOLVES IN EXCESS TO FORM COLOURLESS SOLUTION

Exam Tip

The ammonia or sodium hydroxide solution must be added very slowly. If it is added too quickly and the precipitate is soluble in excess, then you run the risk of missing the formation of the initial precipitate, which dissolves as quickly as it forms if excess solution is added.

Be sure to distinguish between the term "**colourless**" and "**clear**". A solution that loses its colour has become colourless. A clear solution is one that you can see through such as water. Solutions can be clear **and** have colour eg. dilute copper sulphate.

- The flame test is used to identify the metal cations by the colour of the flame they produce
 - Ions from **different** metals produce **different colours**

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Your notes

- Dip the loop of an **unreactive** metal wire such as nichrome or platinum in concentrated acid and then hold it in the blue flame of a Bunsen burner until there is no colour change
 - This is an important step as the test will only work if there is just **one type** of ion present
 - Two or more ions means the colours will mix, making identification erroneous
 - This cleans the wire loop and avoids **contamination**
- A small sample of the compound is placed on an **unreactive** metal wire loop such as nichrome or platinum
- Dip the loop into the solid sample / solution and place it in the edge of the **blue** Bunsen flame
 - Avoid letting the wire get so hot that it glows red otherwise this can be confused with a flame colour

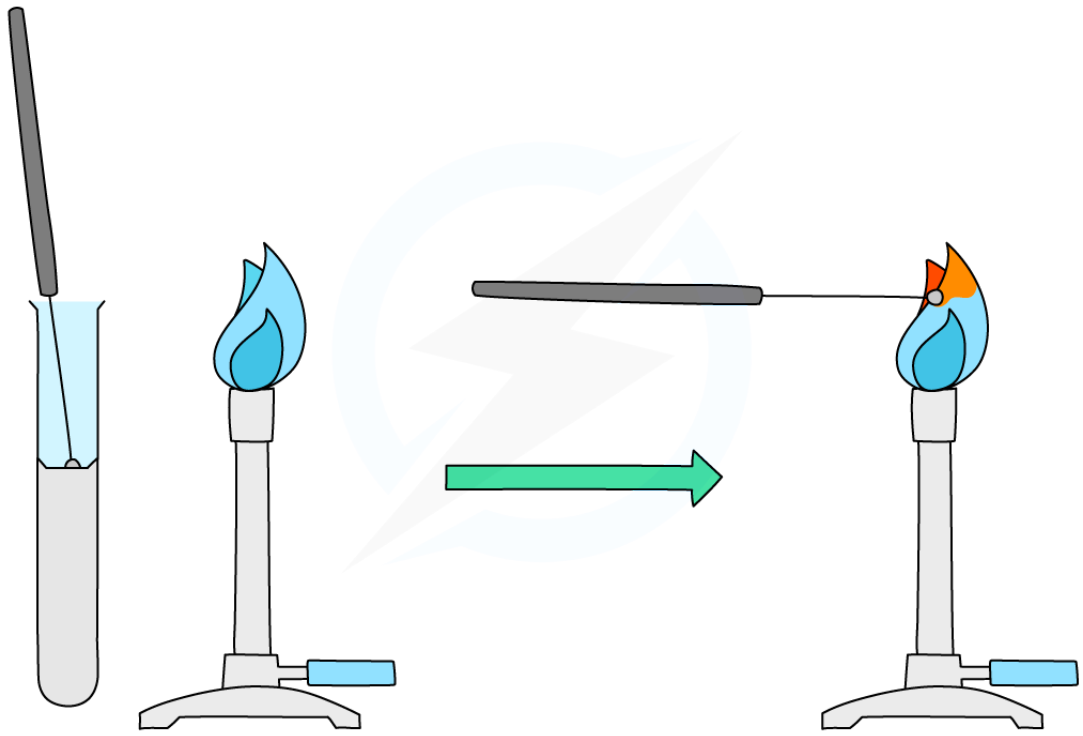

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Diagram showing the technique for carrying out a flame test

- The colour of the flame is observed and used to identify the metal ion present:

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Your notes

Cation	Colour of flame
Li^+	Red
Na^+	Yellow
K^+	Lilac
Ca^{2+}	Orange–red
Cu^{2+}	Blue–green
Ba^{2+}	Apple green

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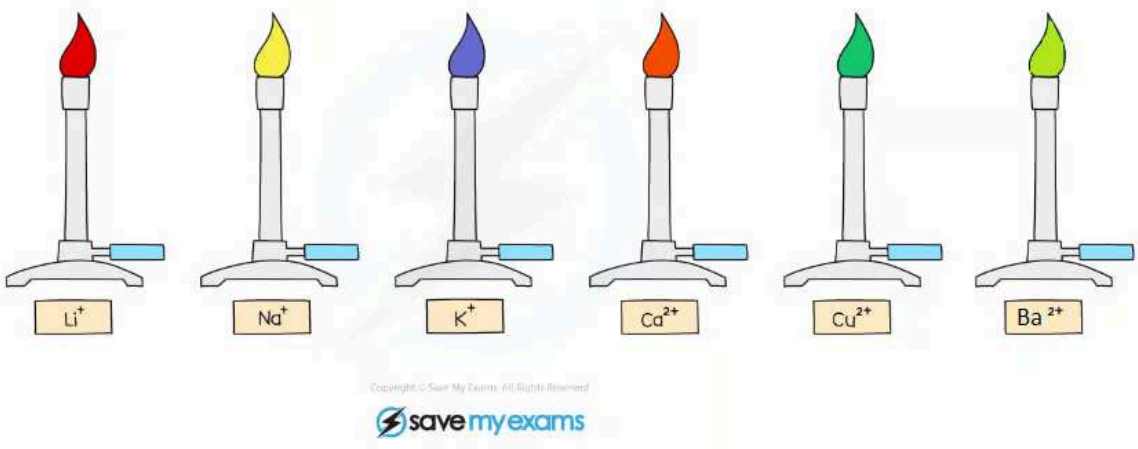


Diagram showing the colours formed in the flame test for metal ions

Exam Tip

The sample needs to be heated strongly, so the Bunsen burner flame should be on a blue flame.

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Identification of Gases

Your notes

Identification of Gases

- Several tests for anions and cations produce **gases** which then need to be tested
- The table below indicates the tests for the gases included in the syllabus

Gas	Appearance of gas	Test	Result
Ammonia (NH_3)	Colourless, pungent smell	Damp red litmus paper	Turns blue
Carbon dioxide (CO_2)	Colourless and odourless	Bubble through limewater	Limewater turns milky/cloudy
Chlorine (Cl_2)	Pale green, choking smell	Damp blue litmus paper	Turns red and quickly is bleached
Hydrogen (H_2)	Colourless and odourless	Hold a lighted splint in mouth of test tube	Burns with a 'squeaky pop' sound
Oxygen (O_2)	Colourless and odourless	Hold a glowing splint	Splint relights
Sulfur dioxide (SO_2)	Colourless, pungent chocking smell	Add to acidified aqueous potassium manganate(VII)	Turns from purple to colourless

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Exam Tip

It is easy to confuse the tests for hydrogen and oxygen. Try to remember that a lig**H**ted splint has a **H** for **H**ydrogen, while a gl**O**wing splint has an **O** for **O**xygen.

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